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NUMERICAL EXERCISES IN CHEMISTRY.

BY

T. HANDS, M.A., F.R.A.S.

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NUMERICAL EXERCISES

IN CHEMISTRY

BY

T. HANDS, M.A., F.R.A.S.

Science Master in Carlisle Grammar School, and late Scholar of Queen's College, Oxford.

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PREFACE.

This book contains 650 Examples, illustrating Chemistry and parts of Physics closely connected with Chemistry. Most of them are original, the rest are taken from various Examination Papers. Each set contains a considerable number of similar examples grouped together, so that the formula by which they are worked may be thoroughly grasped in its different bearings. The bookwork is merely intended to explain each formula and its use, and order of treatment is entirely subordinated to the convenient arranging of examples.

Several difficulties in the text have been ignored, their discussion being beyond the scope of a collection of exercises. For instance, on p. 10, the definition of mass suggests a question as to the sense in which bodies of dissimilar chemical constitution can be said to contain equal quantities of "matter."

For convenience of verification, the numerical data have been so chosen that the answers to nearly all the questions are exact and not approximate.

I shall be glad to receive any suggestions or corrections; my thanks are due to W. Larden, Esq., of Cheltenham College, and to the Rev. Fred. F. Grensted, of the Merchant Taylors' School, Crosby, Liverpool, for valuable assistance of this kind.

T. H.

Carlisle Grammar School,
January, 1884.

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NUMERICAL EXERCISES IN CHEMISTRY.

CHAPTER I.

METRIC SYSTEM OF WEIGHTS AND MEASURES.

1. The nature of the metric system of measurement is fully explained in treatises on Chemistry. It will be sufficient to give here the tables of length, volume, and mass.

TABLE OF LENGTH.

10 millimetres = 1 centimetre.
10 centimetres = 1 decimetre.
10 decimetres = 1 metre.
10 metres = 1 decametre.
10 decametres = 1 hectometre.
10 hectometres = 1 kilometre.

Or again, 1 metre = 100 centimetres = 1000 millimetres. It assists the memory to notice that the names of the submultiples of the metre are derived from Latin numerals, and the names of its multiples from Greek numerals.

MEASURE OF SURFACE.

100 square metres = 1 are.

2. The unit of *Volume* is the litre, *i.e.* a cubic decimetre, or 1000 cubic centimetres. Frequently 1 c.c. is taken as the unit of volume.

TABLE OF VOLUME.

10 millilitres = 1 centilitre.
10 centilitres = 1 decilitre.
&c., &c.

3. The unit of *Mass* is the gram, which is the mass of

1 c.c. of distilled water at 4° C. The standard from which the gram is rigorously determined is a kilogram (1000 grams) of platinum, kept in Paris.

TABLE OF MASS.

10 milligrams = 1 centigram.
10 centigrams = 1 decigram.
&c., &c.

Or again, 1 gram = 100 centigrams = 1000 milligrams.

4. To connect the English system of measurement with the above, we may take

1 foot = 30·48 centimetres.

Ex. (1). *Express in metres the sum of* 6 *kilometres,* 16 *centimetres, and* 8 *millimetres.*

6 kilometres = 6000 metres (multiplying by 1000).
16 centimetres = 0·16 „ (dividing by 100).
8 millimetres = 0·008 „ (dividing by 1000).

∴ sum = 6000·168 metres.

Ex. (2). *How many litres are contained in a rectangular reservoir* 2 *decametres long,* 10 *metres wide, and* 5 *metres deep, and what mass of water would it contain at* 4° *C.?*

Content of reservoir = 20 × 10 × 5 cubic metres.
= 1000 cubic metres.
= 1,000,000 cubic decimetres.
= 1,000,000 litres.

Again, 1 c.c. of water at 4° C. weighs 1 gram.

∴ 1 cubic decimetre, *i.e.* 1 litre, weighs 1000 grams, or 1 kilogram.

∴ mass of water in reservoir = 1,000,000 kilograms.

EXAMPLES I.

1. Express in centimetres—

(1) 7 decametres.
(2) 5 millimetres.
(3) 3 kilometres.
(4) 33 decimetres.
(5) 2 hectometres, 5 metres.
(6) 11 decimetres, 3 millimetres.
(7) 37 metres, 4 decimetres.
(8) ·042 kilometre.
(9) 3·539 metres.

2. Express in metres—

(1) 34 centimetres.
(2) 106 millimetres.
(3) 8 kilometres, 9 decimetres.
(4) 42 decametres, 4 centimetres.
(5) ·27 hectometre.
(6) ·5361 centimetre.

3. Express in litres—

(1) 7000 c.c.
(2) 64 c.c.
(3) 8 cubic metres.
(4) 9500 cubic millimetres.
(5) 5 cubic metres, 5 c.c.
(6) ·003 cubic decametre.
(7) 5·67 cubic decimetres.
(8) 2·002 cubic metres.
(9) 39·5 c.c.

4. Find in centimetres the sum of—

(1) 8 metres, 7 decimetres, 4 millimetres.
(2) 42 decametres, 4 decimetres.
(3) 6 kilometres, 6 metres, 6 centimetres.
(4) 11 metres, 11 decimetres, 11 millimetres.
(5) 1·69 decimetres, 14 millimetres.

5. Reduce to metres, &c.—

(1) 20391 m.m.
(2) 1763·2 cm.
(3) ·007654 kilometre.

6. A rod 15 cm. long is shortened by 26 m.m.; what is the length of the remainder?

7. Into how many pieces, each 5 cm. long, can a metre bar be cut?

8. How many litres are contained in a rectangular trough 1 m. long, 50 cm. wide, and 30 cm. deep?

9. A flask will contain 80 grams of distilled water at 4° C. up to a certain mark; what is the capacity of that portion of it?

10. Three metres of a certain wire weigh ·15 gram; what length of it would weigh a milligram?

11. Find in ares the size of a field 150 metres long, and 60 metres wide.

12. What is the length of the side of a square 36 ares in surface?

13. If a metre scale graduated to millimetres, and a foot scale graduated to eighths of an inch, be placed side by side

with the zeros coincident, which of their other graduations will exactly coincide?

14. A train is travelling 36 miles an hour, what is its velocity in metres per second? (Take a mile as 1·609 kilometre.)

15. Taking a yard as 0·91 metre, find to four places of decimals the value of an acre in ares.

16. Find the value of g in feet per second at a place where its value is 981·456 cm. per second.

CHAPTER II.

THERMOMETRIC SCALES.

1. *Principle of a Thermometer.*

If any quantity of a liquid, such as mercury, be poured into a vessel and the space it occupies observed, it will be found in general that the liquid takes up more room when heat is applied to it, *i.e.* that its volume increases when the temperature is raised. So long as the volume is increasing we infer that the temperature is rising; if the volume remain the same for any period of time we infer that the temperature has not changed during that time; if the volume begin to diminish we infer that the temperature is falling. We may therefore compare the temperatures of two bodies by placing the vessel first in one and then in the other, and noting in each case the space occupied by the mercury. Such a vessel would be a rude thermometer.

2. *Fixed Points.*

To actually *measure* the temperature of a body, we must choose some temperature to which to refer it. If, now, a thermometer be placed in melting ice, the mercury within it takes a definite volume which is seen to be the same each time that the experiment is tried.

Ice, therefore, melts at a definite temperature.

Again, if a thermometer be placed in the steam of water boiling under a pressure of 760 m.m., the mercury takes another definite volume, indicating a definite temperature. These two fixed intensities of heat, the temperatures at which ice melts and water under standard pressure boils, are taken as *fixed points* to which to refer other temperatures. They are called the freezing-point and boiling-point respectively.

Marks being made on the stem to show the height at which the mercury stands in these two cases, the distance between the marks is divided into any number of equal parts called *degrees.* The volume of the mercury at any time can then be recorded by noting the number at the top of the mercury column, and this serves as a measure of the temperature at the time. Any numbers might be written against the two reference marks, the difference between them giving the number of degrees from one mark to the other. Those chosen on the scales in common use are—

	Freezing-point.	Boiling-point.
Fahrenheit's Scale . .	32°	212°
Centigrade " . .	0°	100°
Réaumur's " . .	0°	80°

Consequently the number of degrees between the fixed points is 180, 100, and 80 on the three scales respectively.

3. FORMULÆ OF CONVERSION.

(i.) *Intervals of Temperature.*

Suppose that when a Fahrenheit thermometer is placed in one vessel of water the mercury rests at 40°, and that when it is placed in another the mercury rests at 58°. The difference of temperature indicated is therefore 18° F.; what difference of temperature would a Centigrade thermometer have shown?

The difference between the freezing-point and boiling-point of water is, as we have seen, 180° F. and also 100° C.

$$\therefore\ 180^\circ \text{ F.} = 100^\circ \text{ C.}$$

$$\therefore\ 1^\circ \text{ F.} = \frac{100^\circ}{180} \text{ C.} = \frac{5^\circ}{9} \text{ C.}$$

$$\therefore\ 18^\circ \text{ F.} = 18 \times \frac{5^\circ}{9} \text{ C.} = 10^\circ \text{ C.}$$

Thus the Centigrade thermometer would have risen 10°.

Hence, to find the number of degrees Centigrade in an interval of temperature given in degrees Fahrenheit, we must multiply the given number by 5, and divide by 9. In the same way we see that to find the number of degrees Fahrenheit in an interval of temperature expressed in degrees ntigrade, we must multiply by 9 and divide by 5. Thus, if F is the number of degrees Fahrenheit in the temperature interval, and C the corresponding number of degrees Centigrade, $F = \frac{9}{5} C$.

Similarly, if R is the number of degrees Réaumur in the same interval, we find $F = \frac{9}{4} R$, when we substitute 80 for 100 and R for C in the previous reasoning.

$$\text{Consequently } \frac{F}{9} = \frac{C}{5} = \frac{R}{4} \quad . \; . \; . \; . \; . \; . \; . \; . \quad (1).$$

(ii.) *Temperature Readings.*

Suppose that a Fahrenheit thermometer placed in water indicates 41°; what would a Centigrade thermometer indicate? Place the Fahrenheit thermometer in melting ice; it registers 32°. Place it in the water; it registers 41°; *i.e.* the mercury has risen 9° F. Hence a Centigrade thermometer similarly treated would rise $\frac{5}{9} \times 9°$ C., or 5° C., by what has been already shown. Placed in melting ice it indicates 0° C; placed in the water it indicates 5° more, *i.e.* 5° C. Thus 5° C. is the temperature required. Similarly in other cases. Therefore, to convert a Fahrenheit reading to the Centigrade scale, subtract 32, multiply by 5, and divide by 9. Expressed as a formula, $C = \frac{5}{9} (F - 32)$, C and F here denoting *readings* and not *intervals* of temperature. Conversely, if the Centigrade thermometer indicate 5°, *i.e.* if it rise 5° C. above the freezing-point when placed in the water, a Fahrenheit thermometer would rise 9° F. above the freezing-point, *i.e.* 9° F. above 32° F. Hence the Fahrenheit thermometer would indicate 41°. To convert, then, a Centigrade reading to the Fahrenheit scale, divide by 5, multiply by 9, and add 32.

Expressed as a formula, $F = \frac{9}{5} C + 32$.

Similarly we find that—

$$R = \frac{4}{9} (F - 32), \; F = \frac{9}{4} R + 32.$$

These formulæ may be written—

$$\frac{F-32}{9} = \frac{C}{5} = \frac{R}{4} \quad . \; . \; . \; . \; . \quad (2)$$

where F, C, and R are the readings on the three scales when the three thermometers are at the same temperature.

Ex. *Find the Fahrenheit reading corresponding to* 36°·5 *C.*

Denote it by F, then by (2)

$$\begin{aligned} \frac{F-32}{9} &= \frac{36\cdot5}{5} \\ &= 7\cdot3 \\ \therefore F - 32 &= 7\cdot3 \times 9 \\ &= 65\cdot7 \\ \therefore F &= 32 + 65\cdot7 \\ &= 97\cdot7. \end{aligned}$$

EXAMPLES II.

1. Express the following Centigrade readings on the other two scales:—
(1) 15°. (2) 35°. (3) 50°. (4) −80°. (5) −45°. (6) 21°·5. (7) 33°·5. (8) −6°·5. (9) −30°·5. (10) −272°·9.

2. Transform the following Fahrenheit readings to the other two scales:—
(1) 50°. (2) 68°. (3) 77°. (4) −4°. (5) −22°. (6) 64°·4. (7) 39°·2. (8) −5°·8. (9) −18°·4. (10) −29°·2.

3. Find the readings on the other two scales corresponding to the following Réaumur readings:—
(1) 24°. (2) 56°. (3) 72°. (4) −8°. (5) −44°. (6) 18°·4. (7) 49°·2. (8) −2°·8. (9) −83°·6. (10) −115°·2.

4. Express the following intervals of temperature Fahrenheit on the other two scales:—
(1) 36°. (2) 81°. (3) 45°. (4) 33°. (5) 28°·8.

5. Express the following intervals of temperature Centigrade on the other two scales:—
(1) 25°. (2) 40°. (3) 10°. (4) 55°. (5) 61°·5.

6. A Centigrade thermometer falls 15° on being placed in cold water. How much would a Fahrenheit thermometer have fallen under the same circumstances?

7. What temperature is expressed by the same number—
(*a*) On the Fahrenheit and Centigrade scales?
(*b*) On the Fahrenheit and Réaumur scales?
(*c*) On the Centigrade and Réaumur scales?

8. What temperatures are expressed by integers on all three scales? How many of them are there between 0° C. and 100° C., excluding these temperatures themselves?

9. Compare the sizes of degrees on the same thermometer on the three scales, and find their actual magnitude when the distance between the fixed points is 28·8 cm.

10. Prove that any Fahrenheit reading exceeds the sum of the equivalent Centigrade and Réaumur readings by 32.

11. If any Centigrade reading be divisible by 5, the corresponding Réaumur reading is divisible by 4.

12. Prove that the sum of a Centigrade and the equivalent Réaumur reading is equal to nine times their difference.

13. The sum of a Centigrade and the equivalent Fahrenheit reading is 88; find them.

14. The sum of a Centigrade and the corresponding Fahrenheit reading is 4; find them.

15. The excess of a Fahrenheit over the corresponding Centigrade reading is 36; find them.

16. The excess of a Fahrenheit over the corresponding Centigrade reading is 16; find them.

17. The temperature at which water reaches its maximum density on the Fahrenheit scale is 9·8 times that on the Centigrade scale; find both.

18. Prove that any reading Fahrenheit is equal to nine times the difference of the corresponding Centigrade and Réaumur readings, together with 32.

19. Prove that any Centigrade reading is equal to five times its excess over the corresponding Réaumur reading.

20. If the temperature of the earth rise 1° F. for every 50 ft. that we approach the centre, how many metres will give a rise of 1° C.?

21. The latent heat of steam is 536 thermal units, *i.e.* one gram of steam at 100° C., in condensing into water at 100° C., will raise the temperature of 536 grams of water at zero 1° C. What is the value of the latent heat of steam, using Fahrenheit's scale?

CHAPTER III.

VOLUME, MASS, DENSITY, WEIGHT.

1. THE *volume* of a body is the space it occupies; it is measured in cubic centimetres, cubic inches, &c.

2. Newton defines the *mass* of a body as the quantity of matter it contains. The most convenient unit of mass is the quantity of matter in a cubic centimetre of distilled water at 4° C., viz. a *gram*. The mass of any body is measured in terms of this or some other unit, *e.g.* a pound, by weighing with an ordinary pair of scales.

3. The *density* of a body is the quantity of matter in a unit of volume of it, *i.e.* the mass of unit volume. Call this d, and let M be the total mass, and V the volume of the body. Then since

mass of 1 unit of volume is d
$\therefore$ . . . . 2 units . . . $2d$
. . . 3 . . . . . . $3d$
. . . .
. . . V . . . . . . Vd
$\therefore$ $M = Vd$ . . . . . (1).

Density is usually taken to mean the ratio of the mass of any volume of the substance considered to that of an equal volume of some standard substance. The standard substance for solids and liquids is distilled water at 4° C.; for gases, air or hydrogen at 0° C. and 760 m.m. pressure. The density of a substance in this sense is the same as its specific gravity; if the gram and centimetre be taken as units, the two definitions are identical for solids and liquids, since the mass of unit volume is then the number of grams in a cubic centimetre of the substance.

4. The *weight* of a body is the force with which the earth

"attracts" it. Its weight is the cause of the motion of a falling body; it increases with the mass of the body, and also as the body approaches the earth's centre. Hence the weight of a given body is less at the equator than at the poles. Now it is plain that a pound of matter weighed in a pair of scales at the equator would still weigh a pound at the poles, since any change in the earth's attraction will affect both pans alike. It follows that a pair of scales gives the *mass* of a body and not its *weight*, for the weight varies from place to place. A pound is in fact a *mass* and not a *weight*. The indications of a spring balance, on the contrary, vary with the weight of the body.

Let g denote the acceleration due to gravity, *i.e.* the velocity acquired in each second by a body falling freely *in vacuo* under the influence of the earth's attraction. Then g is the change of the velocity of the body per second, Mg is therefore the change of its momentum. And since W, the weight of the body, is the force producing this change, Newton's Second Law of Motion gives

$$W = Mg = Vdg \quad . \quad . \quad . \quad . \quad . \quad (2),$$

by suitably choosing the units of weight, &c.

The unit of weight is the same as the unit of force, viz. the *dyne*, *i.e.* a force which acting on a gram for one second produces in it a velocity of one centimetre per second.

And since gravity acting on a gram gives it in each second of its motion a velocity of g cm. per second, we see that the weight of a gram is g dynes.

Ex. (1). *What is the density of* 16 *grams of a substance, if its volume is* 8 *c.c.?*

The mass of 1 c.c. is evidently 2 grams, *i.e.* the density of the body is 2, for a given volume of it weighs twice as much as an equal volume of water.

This also follows from equation (1).

$$\text{M, the mass, is 16 grams; V, the volume, 8 c.c.}$$
$$\therefore\ 16 = 8d, \text{ whence } d = 2.$$

Ex. (2). *A flask will hold* 11 *grams of water at* 0° *C.; how much sulphuric acid* (*density* 1·6) *will it hold?*

Since a gram of water at 0° C. occupies 1 c.c. very approximately, the volume of the flask is 11 c.c.

And the mass of 1 c.c. of the acid is 1·6 gram.

$\therefore$ total mass of acid = 11 × 1·6 = 17·6 grams.

Ex. (3). *A body weighs* 7·64 *grams in air, and* 3·64 *grams in water; find its density, neglecting the weight of the air.*

A body suspended in water appears to weigh less than in air by the mass of a quantity of water of the same volume as the body.

This gives us—
mass of equal volume of water = 7·64 − 3·64 = 4 grams.
And by definition—

$$\text{density of body} = \frac{\text{mass of body}}{\text{mass of equal volume of water}} = \frac{7{\cdot}64}{4} = 1{\cdot}91.$$

Ex. (4). 10 *c.c. of alcohol* (*density* 0·79) *are mixed with* 15 *c.c. of water. Find the density of the mixture, neglecting contraction.*

The mass of the mixture is (10 × 0·79 + 15) grams, or 22·9 grams; its volume is 25 c.c.

$$\therefore \text{density of mixture} = \frac{22{\cdot}9}{25} = 0{\cdot}916.$$

EXAMPLES III.

1. Find the density of—
(1) 10 grams of a substance, which occupy 2 c.c.
(2) 13 c.c. of a substance, which weigh 14·3 grams.
(3) 24 c.c. of a substance, their mass being 57·6 grams.

2. Find the volume of—
(1) 65 grams of a substance the density of which is 5.
(2) 28·7 grams of a substance 4·1 times as heavy as water.
(3) 151·42 grams of a substance of density 13·4.

3. Find the mass of—
(1) 11 c.c. of a substance the density of which is 7.
(2) 4·2 c.c. of a substance 2·1 times as heavy as water.
(3) 1·2 c.c. of a substance 4·7 times as heavy as water.

4. If $g = 981$ cm. per second, find in dynes the weight of—

(1) 4 grams of a substance.

(2) 4 c.c. of a substance the density of which is 1·5.

(3) 2·7 grams of a substance.

5. What is the mass of a body the weight of which is 1177·2 dynes?

6. A flask will hold 5 grams of water at 0° C.; how much mercury (density 13·596) will it hold?

7. What is the capacity of a flask which will just contain 203·94 grams of mercury at 0° C.?

8. A spring balance to which is attached a flask full up to a certain mark of ether (specific gravity ·716) indicates a force of 7023·96 dynes; what is the capacity of the flask up to the mark?

9. Prove that the number representing the mass of unit volume is the reciprocal of that representing the volume of unit mass.

10. Find the density of a substance 17 grams of which occupy the same space as 5 grams of water.

11. What would be the height in feet of a water barometer, when the atmosphere supports a column of mercury 762 m.m. high?

12. Find the absolute pressure in dynes per square cm., represented by a barometric height of 75·03 cm., g being 980 and the density of mercury 13·6.

13. If 5 cubic inches of mercury weigh 2·45 lbs., and 2 cubic inches of cast iron weigh 0·52 lbs.; what ratio does the density of mercury bear to that of cast iron?

14. A certain body weighs 12·5 grams in water, 17·3 grams in air, and 13·1 grams in oil. Find the specific gravity of the oil.

15. A piece of antimony weighs 6·3745 grams in air, and 5·4245 grams in water; calculate its density.

16. What height of turpentine (density 0·87) in one arm of a U-tube would balance 29 m.m. of mercury (density 13·596) in the other?

17. 5 parts by volume of sulphuric acid (density 1·8) are mixed with eleven parts of water; find the density of the mixture, neglecting contraction.

18. 2 c.c. of oxygen are mixed with 3 c.c. of hydrogen; find

the density of the mixture when it occupies 7 c.c., taking hydrogen as standard substance, and the density of oxygen 16.

19. Find the specific gravity of an alloy of 12 oz. of copper (sp. gr. 8·8), 3 oz. of tin (sp. gr. 7·4), and 5 oz. of lead (sp. gr. 11·4).

20. A piece of rock crystal (sp. gr. 2·6) weighs 2 oz. in water, and 2·25 oz. in spirit; find the specific gravity of the spirit.

CHAPTER IV.

BOYLE'S LAW.

1. SUPPOSE mercury to be poured into the open end of a bent tube, closed at the other end; the air inside will be confined under a pressure which can be increased at will by pouring in more mercury. Experiment shows that, so long as the temperature is unaltered, if the pressure be doubled the air will occupy only half the space that it did at first; if the pressure be trebled the air will occupy one-third of the original space; and so on. If, on the other hand, the pressure be halved, the air will expand to twice its original volume. The same thing is found to be true of all gases under moderate pressure, and not near their point of liquefaction.

Suppose, for instance, that the pressure on 150 c.c. of gas is diminished from 764 m.m. to 573 m.m.; what will the volume become?

Under 764 m.m. pressure the gas occupies 150 c.c.
∴ . 1 m.m. . . . . . . 150×764 c.c.
∴ . 573 m.m. . . . . . . $\frac{150 \times 764}{573}$ c.c.
i.e. 200 c.c.

2. We can obtain a general formula as follows:—

Suppose the pressure of a given quantity of gas to be p units when the volume is v units, and that on the pressure becoming p' units the volume becomes v' units, the mass and temperature of the gas remaining unaltered throughout.

Then under p units of pressure the volume is v units.
∴ . 1 unit . . . . . pv .
∴ . p' units . . . . . $\frac{pv}{p'}$.

But this is v', by supposition.

$$\therefore \frac{p\,v}{p'} = v'.$$

$$\text{or } p\,v = p'\,v'.$$

Similarly, if on the pressure becoming p'' the volume became v'', we should find that $p\,v = p''\,v''$.

Thus, for any given mass of gas at a fixed temperature, $p\,v$ has always the same value, *i.e.*—

$$p\,v = p'\,v' = k, \text{ a constant } \ldots\ldots \quad (1).$$

This experimental law was first enunciated by Boyle, in England, and later by Mariotte, in France; it is known by the name of either.

If M be the mass and d the density of the gas, $M = V\,d$.

Hence for V we may put $\frac{M}{d}$,

and (1) then becomes $\frac{p\,M}{d} = \frac{p'\,M}{d'} = k$,

$$\text{or } \frac{p}{d} = \frac{p'}{d'} = k'. \quad \ldots\ldots\ldots \quad (2).$$

Ex. 150 *c.c. of gas under* 764 *m.m. pressure expand to* 200 *c.c.; what does the pressure become?*

Denote it by p; then by (1)

$$p \times 200 = 150 \times 764.$$

$$\therefore p = \frac{150 \times 764}{200} \text{ m.m.}$$

$$= 573 \text{ m.m.}$$

3. *Dalton's Law of Gaseous Mixtures.*

If a mixture be made of two or more gases which do not act chemically upon one another, the pressure of the mixture is equal to the sum of the separate pressures of its constituents.

Thus, let v c.c. of one gas at pressure p; v' c.c. of another gas at pressure p', &c., be mixed in a vessel, the volume of which is V. Suppose P the pressure of the mixture.

In the end the gases have all the common volume V, but exert different pressures $\pi, \pi', \ldots.$ which together amount to P.

And by Boyle's Law—

$$p v = \pi V.$$
$$p' v' = \pi' V.$$

. . . .

Therefore by addition—

$$\begin{aligned} p v + p' v' + \ . \ . \ . \ . \ . &= \pi V + \pi' V + \ . \ . \ . \\ &= (\pi + \pi' + \ . \ . \ . \ .) V \\ &= P V. \end{aligned}$$

by what has been already said.

EXAMPLES IV.

1. What volume will 162 c.c. of air under 765 m.m. pressure occupy when the pressure is increased to 810 m.m.?

2. Under what pressure will 118 c.c. of hydrogen at 762 m.m. occupy 127 c.c.?

3. By how much will the volume of 152 c.c. of air at 735 m.m. pressure be diminished, when the pressure is increased by 25 m.m.?

4. 4 c.c. of oxygen under 74 cm. pressure are made to occupy four times the volume; find the pressure the gas then exerts.

5. The volume of 60 c.c. of hydrogen at 845 m.m. is increased by 5 c.c.; find the change of pressure.

6. The pressure of 88 c.c. of gas is increased from 770 m.m. to 968 m.m.; what alteration in bulk occurs?

7. Under what pressure will 216 c.c. of gas at 665 m.m. occupy 190 c.c.?

8. 193 c.c. of air at 828 m.m. pressure are allowed to expand to 207 c.c.; find the new pressure.

9. The pressure of 64 c.c. of gas at 612 m.m. is diminished by 10 cm.; find the space then occupied.

10. How will the pressure of 13 c.c. of gas be affected by allowing it to expand to 39 c.c.?

11. 20·3 c.c. of gas at 836 m.m. expand to 22 c.c.; find the change of pressure.

12. The density of nitrogen at standard pressure and temperature being 14, find its density at 0° C. and 798 m.m.

13. Under what pressure would the density of oxygen at

0° C. be the same as that of nitrogen at standard pressure and temperature, the normal density of oxygen being 16?

14. The density of a certain gas is doubled without change of temperature; what change of pressure would effect this, and how would the volume be affected?

15. What change of pressure would bring the density of hydrogen at standard pressure and temperature to 1·2?

16. Prove that for the same mass of gas the ratio of the decrease of pressure to the increase of volume is proportional to the product of the initial and final volumes.

17. A bubble of air 5 cm. long is confined at the bottom of a narrow bent tube by two columns of mercury, each 10 cm. in height. An additional 10 cm. of mercury is poured into one arm; find the difference in level of the two columns.[1]

18. The section of a barometer tube being constant, find how far the mercury will be depressed on introducing into the space above it a quantity of air which at present atmospheric pressure (760 m.m.) occupies a length of 1 m.m. in the tube, the length of the tube above the mercury being initially 66 m.m., and the level in the cistern being kept constant.

19. Into a vessel of 12 litres capacity are introduced 5 litres of air and 4 litres of oxygen, both at standard pressure. What is the pressure of the mixture if the temperature remain unaltered?

20. Calculate the pressures due respectively to the oxygen and nitrogen of the air when the barometric height is 756 m.m., taking the proportions by volume in which these gases occur in the atmosphere to be 1 : 4.

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[1] *This question is from a paper; the bore must be supposed so small that the air does not escape through the mercury.*

CHAPTER V.

CHARLES' LAW.

1. If heat be applied to gas contained in a vessel closed by a piston, the gas will tend to expand, and will therefore exert a greater pressure on the piston and the sides of the vessel. By raising the piston, and permitting expansion, the additional pressure may be relieved, and the temperature of the gas raised while its pressure remains the same. Suppose, now, that heat is applied to 273 c.c. of gas at 0° C., and that the temperature rises to 1° C., while the pressure remains unaltered. The gas will then be found to occupy 274 c.c., or 1 c.c. more than at 0° C.; it has expanded $\frac{1}{273}$ of its volume at 0° C. If heated to 2° C. it will occupy 275 c.c., or 2 c.c. more than at 0° C., and so on. If 273 c.c. of gas at 0° C. be heated to t° C., the volume will become $(273 + t)$ c.c., provided that the pressure and mass are kept the same throughout the process.

It follows that an amount of gas which occupies 1 c.c. at 0° C. will occupy $\frac{273 + t}{273}$ c.c. at t° C. under the same pressure. Therefore V_0 c.c. at 0° C. would expand to $V_0 \times \frac{273 + t}{273}$ c.c. at t° C.

Call this volume V, then—

$$V = \frac{V_0 (273 + t)}{273}$$

$$\text{or } \frac{V}{273 + t} = \frac{V_0}{273} \quad . \quad . \quad . \quad . \quad . \quad . \quad (1).$$

This formula has been deduced from results obtained first by Charles, and afterwards stated independently by Dalton and Gay-Lussac. Gay-Lussac showed that if V is the volume at t° C. of a quantity of gas which occupies V_0 c.c. at 0° C., under the same pressure,

$$V = V_0 (1 + a t),$$

$$\text{where } a = 0{\cdot}003665 = \frac{1}{272{\cdot}85} = \frac{1}{273}, \text{ nearly.}$$

$$\therefore V = V_0 \left(1 + \frac{t}{273}\right)$$

$$= \frac{V_0}{273} (273 + t),$$

which is the same as (1).

If $t = -273^\circ$ C., this leads to the curious result—

$$V = 0,$$

i.e. it implies that if the temperature could be reduced to -273° C. the gas would shrink into nothing. Temperature reckoned from -273° C. as zero is called *absolute temperature.* If T be the absolute temperature corresponding to t° C., and T_0 that corresponding to 0° C.,

$$T = 273 + t, \quad T_0 = 273.$$

(1) may then be written—

$$\frac{V}{T} = \frac{V_0}{T_0}.$$

Thus the volume of the gas is proportional to the absolute temperature, the pressure being constant. If p be the pressure, we have, by multiplication,

$$\frac{p\,V}{T} = \frac{p\,V_0}{T_0} \quad . \; . \; . \; . \; . \; . \; . \; . \; . \; . \quad (2).$$

Again, suppose the pressure to have been such that the gas at 0° C. occupied V_0' c.c., and that when the temperature became T'° of the absolute scale the volume became V'.

Then, as before,

$$\frac{V'}{T'} = \frac{V_0'}{T_0}.$$

If p' is the pressure in this case,

$$\frac{p'\,V'}{T'} = \frac{p'\,V_0'}{T_0} \quad . \; . \; . \; . \; . \quad (3).$$

But by Boyle's Law, $p\ V_0 = p'\ V_0'$; since the mass and temperature of the gas are the same in each case.

Hence, comparing (2) and (3), we have—

$$\frac{p\ V}{T} = \frac{p'\ V'}{T'},$$

$$\text{or } \frac{p\ V}{273 + t} = \frac{p'\ V'}{273 + t'} \quad . \quad . \quad . \quad (4).$$

This is the law known by either of the names mentioned above.

Ex. 261 *c.c. of gas are collected at* 17° *C., and under a pressure of* 762 *m.m. At what temperature will the volume become* 254 *c.c. under a pressure of* 756 *m.m.?*

Denote the temperature by t; then by (4)—

$$\frac{756 \times 254}{273 + t} = \frac{762 \times 261}{273 + 17}$$

$$\therefore\ 273 + t = \frac{756 \times 254 \times 290}{762 \times 261}$$

$$= 280.$$

$$\therefore\ t = 7^\circ \text{ C.}$$

2. If M is the mass and d the density of the gas, we know that $M = V\,d$.

Hence $V = \frac{M}{d}$, whence (4) may be written—

$$\frac{p}{(273 + t)\,d} = \frac{p'}{(273 + t')\,d'} \quad . \quad . \quad . \quad (5).$$

Ex. *Find the density of sulphuretted hydrogen gas* (H_2S) *at* 27° *C. and* 760 *m.m.*

At 0° C. and 760 m.m. the density of a gas referred to hydrogen is half its molecular weight.

Hence, the density of H_2S at 0° C. and 760 m.m. is $\frac{1}{2}(2 + 32)$, or 17 (Chap. IX.).

Then by (5), since $p = p' = 760$ m.m.; $d = 17$; $t = 0^\circ$ C.; $t' = 27^\circ$ C.,

we have—

$$\frac{760}{273 \times 17} = \frac{760}{(273 + 27)\,d'}$$

$$\therefore\ d' = \frac{273 \times 17}{300}$$

$$= 15{\cdot}47.$$

3. It is possible, but more tedious, to work examples on Charles' Law without introducing an equation. As an instance, let us take Ex. (1) of paragraph (1). Here the volume changes, owing to changes in pressure and temperature; we first get rid of the former, and then deal with the latter from the considerations on page 20.

Under a pressure of 756 m.m. the gas occupies 254 c.c.
∴ „ „ 1 m.m. „ „ 254×756 c.c.
∴ „ „ 762 m.m. „ „ $\frac{254 \times 756}{762}$ c.c.
or 252 c.c.

Thus, by reduction of *temperature only*, 261 c.c. at 17° C. are to contract to 252 c.c. We have to find at what temperature this will be the case.

Now we have shown that when the pressure is constant, the volume is proportional to the absolute temperature. Thus, at constant pressure,
when the volume is 261 c.c., the absolute temperature is 290°.
∴ . . . . . 1 c.c. . . . . . $\frac{290^\circ}{261}$
∴ . . . . . 252 c.c. . . . . . $\frac{290 \times 252^\circ}{261}$
or 280°.

$$\therefore T = 280^\circ$$
$$\therefore t = 7^\circ \text{ C.}$$

Formula (4) might have been obtained in a similar manner.

EXAMPLES V.

1. 39 c.c. of gas are collected at 47° C., and under 320 m.m. pressure; if the pressure be further diminished to 195 m.m., and the temperature raised to 112° C., what will the volume become?

2. 57 c.c. of gas, originally at 16° C. and 272 m.m., occupy 38 c.c. at 33° C.; find the pressure.

3. The temperature of a certain quantity of gas is brought

from 70° C. to 28° C., and the pressure from 735 m.m. to 516 m.m. The volume being then 185 c.c., find the original volume.

4. At what temperature will 182 c.c. of air now at 0° C. occupy 186 c.c., the pressure being unaltered?

5. Find the change of pressure when 92 c.c. of air at 23° C. and 502 m.m. are brought to occupy 69 c.c. at 60° C.

6. At 54°·6 C. and 366 m.m. a certain quantity of gas occupies 12 c.c. more than at zero, when the pressure was 427 m.m. Find the original volume.

7. 45 c.c. of gas at 24° C. and 572 m.m. are allowed to expand 7 c.c., the pressure then falling 36 m.m. Find the final temperature.

8. 82½ c.c. of gas at −9° C. exert a pressure of 376 m.m.; if the temperature be reduced 22° C., and the volume 27½ c.c., find the resulting pressure.

9. Prove that, if temperature be reckoned from −273° C. as zero, the volume of a gas is proportional to its temperature as long as the pressure is unaltered, and the pressure proportional to the temperature as long as the volume is unaltered.

10. What change (*a*) of pressure only, (*b*) of volume only, will raise the temperature of 82 c.c. of gas at 14° C. and 369 m.m. by 7° C.?

11. Prove that if either pressure or volume be constant, the change of the other is proportional to the change of temperature.

12. If the volume of a gas be 18 litres under a pressure of 450 m.m. and at 15° C., what will its volume be under a pressure of 350 m.m. and at 70° C.?

13. If a cubic foot of dry air weigh 540 grains when the thermometer stands at 14° C. and the barometer at 30 inches, find the mass of an equal volume of dry air at 16°·8 C., when the barometer stands at 29·9 inches.

14. A certain mass of air exerts at 0° C. a pressure of 73·437 cm., and at t° C., its volume being constant, a pressure of 96·033 cm. Find t; given that the increment of pressure between 0° C. and 100° C. is $\frac{100}{273} \times$ pressure at 0° C.

15. At standard pressure and temperature the density of a gas relatively to hydrogen is half its molecular weight;

calculate the densities of the following gases at 15° C. (see Chap. IX.) :—

(1) Oxygen. (2) Sulphur Dioxide (SO_2). (3) Marsh gas (CH_4). (4) Nitrogen Dioxide (NO).

16. The mass of a litre of hydrogen at 0° C. and 760 m.m. being 0·0896 gram, find the mass of a litre of the following gases (*a*) at 0° C., (*b*) at 15° C. :—

(1) Oxygen. (2) Nitrogen. (3) Nitrogen Monoxide (N_2O). (4) Nitrogen Dioxide (NO). (5) Ammonia (NH_3) (6) Carbon Monoxide (CO). (7) Carbon Dioxide (CO_2). (8) Marsh gas (CH_4). (9) Chlorine. (10) Sulphur Dioxide (SO_2).

CHAPTER VI.

GRAHAM'S LAW OF DIFFUSION.

1. If a glass tube closed at one end by a porous plug be filled with hydrogen, and the open end placed under water, the water will be seen to rise in the tube. The reason is that the hydrogen is passing out through the plug faster than air passes in, owing to its small density. Graham found that the rates of diffusion of different gases through such a plug, measured by the *volumes* passing through in a given time, were *inversely* proportional to the square roots of their densities.

It follows that if V and V' are the volumes of two gases that pass through in the same time, and d and d' their densities,

$$\frac{V}{V'} = \sqrt{\frac{d'}{d}} \quad . \quad . \quad . \quad . \quad . \quad . \quad . \quad (1).$$

2. If M and M' are the masses of the two gases, we know that $M = V\,d$, $M' = V'\,d'$, whence $V = \frac{M}{d}$ and $V' = \frac{M'}{d'}$.

Thus (1) may be written—

$$\frac{M\,d'}{M'\,d} = \sqrt{\frac{d'}{d}}.$$

$$\therefore \frac{M}{M'} = \sqrt{\frac{d'}{d}} \times \frac{d}{d'}.$$

$$= \sqrt{\frac{d}{d'}} \quad . \quad . \quad . \quad . \quad . \quad (2).$$

Hence the rate of diffusion, measured by the *mass* passing across in a given time, is *directly* proportional to the square root of the density of the gas.

For instance, if a gas were four times as dense as another,

$$\frac{d'}{d} = 4.$$

$$\therefore \frac{V}{V'} = 2, \text{ but } \frac{M}{M'} = \frac{1}{2},$$

i.e. in a given time twice as great a volume of the one as of the other would pass through the plug, but only half as great a mass.

EXAMPLES VI.

1. If 3 c.c. of hydrogen diffuse across a porous plate in 90 minutes, what volume of oxygen would pass through it in two hours?
2. Compare the densities of two gases, equal volumes of which pass through a porous plug in one hour and four hours respectively.
3. What mass of hydrogen would diffuse across a porous plug in the same time as 20 grams of oxygen at the same temperature and pressure?
4. Compare the rates of diffusion of marsh gas and sulphur dioxide.
5. Compare the rates of diffusion of nitrogen and carbon monoxide.
6. What volume of carbon dioxide would diffuse across a stucco plate in the same time as 2·464 grams of nitrogen monoxide, both being at standard temperature and pressure?

CHAPTER VII.

SPECIFIC HEAT.

1. EXPERIMENT shows that different substances require different quantities of heat to alter the temperature of equal masses to the same extent; thus it takes twice as much heat to raise the temperature of a gram of iron 1° as to raise that of a gram of tin 1°. The quantity of heat required to raise the temperature of a given mass of a substance one degree compared with that required to raise the temperature of an equal mass of water one degree is called the *specific heat* of the substance.

2. To compare one quantity of heat with another, we must choose some unit in terms of which to measure them. A convenient unit is the *calorie, i.e.* the amount of heat that will raise the temperature of one gram of water near 0° C. by one degree Centigrade. With this as unit, the specific heat of a substance is the number of calories that will raise the temperature of one gram of it 1° C., the specific heat of water being then unity.

Ex. 10 *grams of a certain metal are heated to* 100° *C., and placed in* 75 *grams of water at* 4° *C. The final temperature of the water is* 10° *C.; find the specific heat of the metal.*

Denote it by c. Then 1 gram of the metal loses c calories in cooling 1° C.; therefore 10 grams lose $10\,c$ calories in cooling 1° C., and in cooling from 100° C. to 10° C., *i.e.* 90° C., they will lose $10\,c \times 90$ calories.

Similarly, 75 grams of water gain 75×6 calories when their temperature rises from 4° C. to 10° C., *i.e.* 6° C. But plainly the heat gained by the water is furnished by the metal, so that these quantities are equal.

$$\therefore 10 \times 90c = 75 \times 6.$$

$$\therefore c = \frac{75 \times 6}{10 \times 90} = 0{\cdot}5.$$

This method of measuring specific heat is called the

3. *Method of Mixture.*

Let M be the mass of a body in grams, and let it be heated to T°. Place it in m grams of water at $t°$ (a lower temperature than T°), contained in a vessel called a calorimeter. Part of the heat of the body goes to raise the temperature of the water, to $\theta°$, say; part raises that of the calorimeter to $\theta°$. Suppose this latter would have raised the temperature of p grams of water from $t°$ to $\theta°$; then p is called the *water-value* of the calorimeter. As in the previous example,

$$\text{Heat lost by the body} = \mathrm{M}\,c\,(\mathrm{T} - \theta).$$

Heat gained by the water and calorimeter $= (m + p)\,(\theta - t)$. But these must be equal, since the former supplies the latter.

$$\therefore \mathrm{M}\,c\,(\mathrm{T} - \theta) = (m + p)\,(\theta - t) \quad . \quad . \quad . \quad . \quad . \quad . \quad (1).$$

This equation determines c.

If the method of heating adopted be to place the body in boiling water, a certain amount of water, P grams, say, at T°, will be carried over when the body is transferred to the calorimeter. This must be allowed for, and (1) becomes

$$(\mathrm{M}\,c + \mathrm{P})\,(\mathrm{T} - \theta) = (m + p)\,(\theta - t) \quad . \quad . \quad . \quad . \quad . \quad (2).$$

We find P by weighing the calorimeter with its contents before and after the transfer.

It should be observed that in all that has preceded we have supposed no heat to be lost to surrounding bodies by radiation, conduction, or convection.

4. *Bunsen's Ice-Calorimeter.*

In this instrument, the heat furnished by the body of which the specific heat is required is employed to melt ice at 0° C. The water formed occupying a less volume than the ice, the quantity of ice melted can be estimated by observing the motion of a mercury column indicating the change of volume. And since the quantity of ice melted depends on the amount of heat given up by the body in cooling to 0° C., the motion of the mercury serves to measure the amount of heat given up. This, when compared with the heat given up by an equal mass of water under similar circumstances, determines the specific heat of the substance.

If X is the distance moved through by the mercury column when M grams of a substance, the specific heat of which is c, are cooled from T° C. to 0° C.; x, m, t, similar quantities for water; evidently

$$\frac{X}{x} = \frac{M c T}{m t} \quad . \quad . \quad . \quad . \quad . \quad . \quad (3).$$

which determines c.

Suppose, for instance, that when 5 grams of silver at 80° C. are put into the calorimeter the mercury column retreats through 16 divisions, and through 64 divisions when 3 grams of water at 32° C. are employed.

Then, if c is the specific heat of silver,

$$\frac{16}{64} = \frac{5 \times c \times 80}{3 \times 32}$$

$$\therefore c = \frac{16 \times 3 \times 32}{64 \times 5 \times 80}$$

$$= \frac{3}{50}$$

$$= 0{\cdot}06.$$

5. It is worth while to remember that water has a greater specific heat than any other substance, so that the specific heat of any other substance is a quantity less than 1.

The following approximate specific heats will be required in the solution of examples in which the values are not given.

	Specific Heats.
Copper	0·0949
Mercury	0·033
Bismuth	0·03
Antimony	0·051
Lead	0·03
Tin	0·05
Iron	0·11

EXAMPLES VII.

1. 45 grams of water at 49° C. are poured into 87 grams at 5° C.; find the final temperature.

2. 18 oz. of water at 63° F. are mixed with 36 oz. at 0° F.; what is the temperature of the mixture?

3. How many grams of water at 62° must be mixed with 54 grams at 12° to bring the temperature of the mixture to 17°?

4. What must be the temperature of 6 grams of water, in order that when mixed with 24 grams at 12° the temperature of the whole may be 11°?

5. A copper ball, weighing 100 grams, is heated to 64°, and immersed in 189·8 grams of water at 22°; find the final temperature.

6. Calculate the specific heat of iron when 12 grams of it at 100° raise the temperature of 66 grams of water from 13°·3 to 15°.

7. 33 grams of copper at 50° are placed in 189·8 grams of mercury at 20°; find the resulting temperature.

8. To what temperature would 15 grams of antimony at zero be raised by a quantity of heat sufficient to bring 42 grams of bismuth from 0° to 34°?

9. 196·7 grams of water and 100 grams of mercury are raised to the boiling-point of the former, and then poured into 300 grams of water at 0° C. Find the temperature of the mixture.

10. In Bunsen's ice-calorimeter, 10·5 grams of water at 34°·2 caused a retreat of the mercury column through 75 divisions, and 61·6 grams of copper at 72° caused a retreat through 88 divisions. Calculate the specific heat of copper.

11. Into a calorimeter containing 233·65 grams of water at 10°·9 C. are thrown 192 grams of tin at 97°·4 C., and the final temperature of the mixture is 14°·1 C. Find the specific heat of tin, the water-equivalent of the calorimeter being 16·25 grams.

12. 80 grams of bismuth heated to 99°·9 are thrown into a calorimeter, containing 117·8 grams of water, a stirrer, and a thermometer at 12°·69. The final temperature is 14°·4;

find the specific heat of bismuth, taking 5·4 grams as the water-value of the calorimeter, stirrer, and thermometer.

13. Calculate the specific heat of tin from the following data:—

Mass of tin at 98°, 51·2 grams in small pieces.
Water-value of wire basket holding it, 1·6 grams.
Mass of water carried over in basket, 6·84 grams.
Water-value of calorimeter, 9·97 grams.
Mass of water in calorimeter, at 17° C., 181·53 grams.
Final temperature, 21°·4 C.

CHAPTER VIII.

LATENT HEAT.

1. If a Centigrade thermometer be placed in a flask of water and heat applied, the temperature of the water will steadily rise to 100°, and then the water will boil. If the supply of heat be continued, the water will continue boiling, but the thermometer will still register 100°. A thermometer, then, fails to indicate that any heat is passing into the water after boiling has once begun, hence such heat has been called *latent, i.e.* hidden. What really becomes of it will be more fully explained in Chapter XIII.; it will there be shown that it is spent in the work of vaporization, and is restored when this work is undone by condensing into water the steam formed.

Similarly when ice is melted, heat is spent in producing the change from the solid to the liquid state, and a Centigrade thermometer, placed in melting ice, indicates 0° till all the ice is melted.

2. The *latent heat* of a substance is the amount of heat which disappears in the manner just explained when unit mass of the substance changes its state without change of temperature. It is measured by means of a calorimeter in much the same way as specific heat.

Ex. 9 *grams of ice at* 0° *C. are dissolved in* 13 *grams of water at* 50° *C. contained in a calorimeter, of which the water-value is* 4 *grams. The final temperature is* 5° *C.; calculate the latent heat of ice.*

Denote it by L; then 9 L calories is the quantity of heat required to convert the 9 grams of ice at 0° C. into water at 0° C., and 9 × 5 calories are needed to raise the temperature of the water formed to 5° C.

Hence—

Heat gained by the ice and by the water formed $= 9\,\mathrm{L} + 45$ calories.

Heat lost by the calorimeter and the water in it $= (13 + 4)(50 - 5)$ calories.

$= 17 \times 45$ calories.

$$\therefore 9\,\mathrm{L} + 45 = 17 \times 45.$$

$$\therefore \mathrm{L} = 80.$$

3. As a more general case, suppose L to be the latent heat of a substance in the solid state, and let m grams of it at t° C. be dissolved in M grams of water at T° C. (where T is higher than t), contained in a calorimeter of which the water value is P. Let θ° C. be the final temperature, and c the specific heat of the substance in the liquid state.

Then $m\,\mathrm{L}$ calories is the heat needed to liquefy m grams of the substance at θ° C. without change of temperature, and $m\,c\,(\theta - t)$ calories the heat required to raise the temperature of the liquid formed from t° C. to θ° C.

$\therefore$ heat absorbed by substance $= m\,\mathrm{L} + m\,c\,(\theta - t)$.

Also, heat lost by water and calorimeter $= (\mathrm{M} + \mathrm{P})\,(\mathrm{T} - \theta)$.

$\therefore m\,\mathrm{L} + m\,c\,(\theta - t) = (\mathrm{M} + \mathrm{P})\,(\mathrm{T} - \theta)$,

since the heat absorbed by the substance is supplied by the water and the calorimeter.

A similar equation to determine L can be found in other cases.

4. Throughout the following examples, when temperature is expressed on the Centigrade scale, the latent heat of ice is assumed to be 80, and that of steam 536.

EXAMPLES VIII.

1. Find the temperature of equilibrium, when 4 grams of snow at zero are dissolved in 12 grams of water at 40° C.

2. How much snow at 0° C. must be dissolved in 20 grams of water at 25° C. to reduce it to the temperature of maximum density?

3. How much ice at zero must be melted in 36 grams of

water at 76° C. to lower the temperature 12° C., assuming no heat lost by radiation?

4. Find the final temperature, when 7 grams of ice at − 6° C. are put into 48 grams of water at 27° C. (Specific heat of ice = 0·5.)

5. Find the temperature to which 150 oz. of water at 4° C. would be raised by condensing in it 8 oz. of steam at 100° C.

6. What would be the result of surrounding 79·5 grams of ice at 0° C. with 10 grams of steam at 100° C.?

7. 30·6 grams of ice at zero are put into 184·88 grams of water at 20°·2 C., contained in a calorimeter whose water value is 9·97. The temperature of equilibrium is 6°·6 C.; find the latent heat of ice.

8. One pound of ice at 0° C. is placed in two pounds of water at 0° C.; find how much steam at 100° C. will melt the ice and raise the temperature of the mixture to 11° C.

9. An oz. of steam at 100° C. is passed into 9 oz. of water at 15° C. The resulting temperature of the 10 oz. of water is 77° C.; find the latent heat of steam.

10. A beaker contains 1340 grams of water at 10° C.; this water is heated by introducing steam at 100° C. What will be the mass of water in the beaker when the steam ceases to be condensed?

11. 11 grams of tin at 235° C. (its melting point) are poured into 100 grams of mercury at 58° C. The heat of liquefaction of tin being 14·25, find the final temperature.

12. Into 50 grams of mercury at 62° C. are poured 5·5 grams of molten lead at 334° C. (its melting point). The final temperature is 103° C.; find the heat of liquefaction of lead.

13. Find the latent heat of ice, using Fahrenheit's scale of temperature.

CHAPTER IX.

CHEMICAL EQUATIONS.

1. *Calculation of Molecular Weights.*

On examining a list of elements such as that on p. 72, it is seen that to each element is affixed a symbol and a number. The first is an abbreviation of the name of the element; the second expresses the ratio of the weight of an atom of it to that of an atom of hydrogen. Hydrogen being the lightest body known, if the weight of an atom of it be denoted by 1, the numbers denoting the weights of atoms of other substances will be greater than 1. Thus the number 39 placed against potassium signifies that an atom of potassium is 39 times as heavy as an atom of hydrogen.

A molecule of a substance, simple or compound, is the smallest portion of it that can exist uncombined. As a rule, an atom of an element cannot exist alone in the free state; oxygen, for instance, is made up of molecules consisting of pairs of atoms. Each of these molecules is represented by the formula O_2. Again, $K\,Cl\,O_3$ represents a molecule of potassium chlorate, consisting of one atom of potassium (K), one of chlorine (Cl), and three of oxygen (O_3). Now an atom of potassium is 39 times as heavy as one of hydrogen, and one of chlorine 35·5 times as heavy. Again, an atom of oxygen is 16 times as heavy as one of hydrogen; three atoms of oxygen are therefore 48 times as heavy. But the weight of a compound is plainly the sum of the weights of its component parts. Hence denoting the weight of a hydrogen atom by 1, the weight of a molecule of potassium chlorate will be

$$39 + 35{\cdot}5 + 48,$$

or 122·5.

Ex. *Find the molecular weight of zinc nitrate,* $Zn\,(N\,O_3)_2$.

We must first find the molecular weight of the group of atoms, $N O_3$, and after doubling it, add it to the weight of an atom of zinc.

Using symbols to abbreviate the work, we have

$$\begin{array}{rcl} N & = & 14 \\ O_3 & = & 48 \text{ because } O = 16. \\ \hline \therefore \underline{\underline{N O_3}} & = & \underline{\underline{62.}} \\ \therefore (N O_3)_2 & = & 124 \\ \text{and} \quad Zn & = & 65 \\ \hline \therefore \underline{\underline{Zn\,(N O_3)_2}} & = & \underline{\underline{189}} \end{array}$$

Thus the molecular weight of zinc nitrate is 189 times the weight of an atom of hydrogen.

The following molecular weights are similarly calculated, and will serve as exercises for the student:—

		Molecular Weight.
(1) Potassium nitrate	$K N O_3$	101
(2) Zinc chloride	$Zn\ Cl_2$	136
(3) Lead nitrate	$Pb\ (N O_3)_2$	331
(4) Ammonium carbonate	$(H_4N)_2C\ O_3$	96
(5) Lead acetate	$Pb\ (C_2H_3O_2)_2$	325

2. *Chemical Equations.*

The change that takes place when potassium chlorate ($K\ Cl\ O_3$) is decomposed by heat into oxygen and potassium chloride ($K\ Cl$) may be thus represented:—

$$\begin{array}{ccccc} 2\ K Cl O_3 & = & 3\ O_2 & + & 2\ K\ Cl \\ 245 & & 96 & & 149 \end{array}$$

Note that all the atoms on one side of the equation reappear on the other (as we should expect, since the substances on the right-hand side are obtained from that on the left), but that they are differently arranged. This equation tells us that two molecules of potassium chlorate, weighing 245 hydrogen-units, yield three molecules of oxygen and two of potassium chloride, weighing respectively 96 and 149 hydrogen-units. The change in question will always take place in these proportions by weight, so that 245 grams, oz., &c., of potassium chlorate will always furnish on decomposition 96 grams, oz., &c., of oxygen and 149 of potassium chloride. Knowing this, we can easily calculate how much of the salt will serve to prepare any quantity of gas required, &c.

Ex. 1. *What mass of oxygen can be obtained from* 147 *grams of potassium chlorate?*

We know that 245 grams of potassium chlorate will furnish 96 grams of oxygen.

$$\therefore \quad 1 \text{ gram} \quad ,, \quad \frac{96}{245} \quad ,,$$

$$\therefore \; 147 \text{ grams} \quad ,, \quad \frac{96 \times 147}{245} \quad ,,$$

i.e. 57·6 grams of oxygen.

Ex. 2. *What quantity of potassium chlorate will furnish* 9 *litres of oxygen at* 0° *C. and* 760 *m.m.?*

Since a litre of hydrogen at 0° C. and 760 m.m. weighs 0·0896 gram, and the density of oxygen referred to hydrogen is 16, the mass of 9 litres of oxygen will be

$$9 \times 0{\cdot}0896 \times 16 \text{ grams.}$$

Consequently the mass of potassium chlorate required is given by the proportion—

$$96 : 9 \times 0{\cdot}0896 \times 16 :: 245 : x.$$

$$\therefore x = \frac{9 \times 0{\cdot}0896 \times 16 \times 245}{96}$$

$$= 32{\cdot}928 \text{ grams.}$$

Ex. 3. *How much potassium chlorate will furnish* 116 *litres of oxygen at* 17° *C. and* 950 *m.m.?*

If V is the volume at 0° C. and 760 m.m. of this amount of gas,

$$V = \frac{116 \times 950 \times 273}{290 \times 760} = 136{\cdot}5 \text{ litres. (Chap. V.)}$$

Whence, as in example (2),

$$96 : 136{\cdot}5 \times 0{\cdot}0896 \times 16 :: 245 : x.$$

$$\therefore x = \frac{136{\cdot}5 \times 0{\cdot}0896 \times 16 \times 245}{96}$$

$$= 499{\cdot}408 \text{ grams.}$$

3. Since 1 litre of hydrogen at 0° C. and 760 m.m. weighs 0·0896 gram, 1 gram of hydrogen will occupy $\frac{1}{0{\cdot}0896}$ litres or very approximately 11·16 litres.[1] Again, 16 grams of oxygen occupy the same space as 1 gram of hydrogen, consequently 16 grams of oxygen occupy about 11·16 litres. And generally, if *m* be the density referred to hydrogen of any substance in the gaseous state, *m* grams of that substance in the gaseous state will occupy approximately 11·16 litres at 0° C. and 760 m.m. The density of a gas is half its molecular weight; the following values should be noticed:—

	Vapour Density.
Phosphorus	62
Arsenic	150
Mercury	100

Hence,—

(1) Given the mass of a substance in grams, to find its volume in litres divide by its density and multiply by 11·16.

(2) To find the mass in grams of a volume given in litres, divide by 11·16 and multiply by the density of the substance in the gaseous state.

We thus get results nearly correct without the labour of exact calculation; of course in all cases the volume must be reduced to standard pressure and temperature. The answers to the examples of this chapter are not calculated by these rules, and will differ slightly from those obtained by using them.

4. In calculating gaseous volumes much trouble is saved by remembering that a molecule of one gas occupies the same space as a molecule of another at the same pressure and temperature. Thus, H_2, O_2, CO_2, NO, . . . represent *equal* volumes of the gases in question under similar conditions of pressure and temperature.

For instance, the equation—

$$H_2 + Cl_2 = 2\,HCl,$$

tells us without any calculation that 2 volumes (*e.g.* 2 litres,

---

[1] If we take the more accurate value of the mass of a litre of hydrogen, viz. 0·089586 gram, we get for the volume of 1 gram of hydrogen 11·162458 . . . litres.

2 c.c., 2 cubic feet, &c.) of hydrogen combine with an equal volume of chlorine to form a volume of H Cl double that of either the hydrogen or chlorine.

5. Throughout the following examples the temperature is supposed to be 0° C. and the pressure 760 m.m. unless otherwise stated. The numbers in square brackets refer to the equations on p. 73.

EXAMPLES IX.

Oxygen.

1. Find how much potassium chlorate is needed to supply the following quantities of oxygen, and the amount of potassium chloride formed:—

(1) 192 grams. (2) 17·28 oz. (3) 21·12 grams. (4) 28·8 grams. (5) 73·92 grams. [1.]

2. Find how much potassium chlorate will yield the following quantities of oxygen:—

(1) 15 litres. (2) 1875 c.c. (3) 2250 c.c. (4) 55 litres. [1.]

3. Find how much mercuric oxide would furnish the following quantities of oxygen:—

(1) 2 grams. (2) 7 grams. (3) 5 litres. (4) 2·6 oz. (5) 6250 c.c. [2.]

4. What mass of oxygen will be obtained by heating to redness the following quantities of manganese dioxide?

(1) 65¼ lbs. (2) 195·75 grams. (3) 182·7 grams. (4) 287·1 oz. (5) 548·1 grams. [3.]

5. In the three methods of preparing oxygen, how much (1) mercury, (2) potassium chloride, (3) red oxide of manganese, is produced simultaneously with 48 grams of the gas?

6. How many c.c. of oxygen can be got by heating 15·12 grams of mercuric oxide? [2.]

7. A globe of capacity 1125 c.c. is to be filled with oxygen at 21° C. and standard pressure; how much mercuric oxide will theoretically suffice? [2.]

8. A certain quantity of potassium chlorate contains 26 grams of potassium; what mass of oxygen would be got by heating it? [1.]

9. How much copper oxide would be formed by burning 25·4 grams of copper wire in excess of oxygen? [4.]

10. Assuming a square centimetre of leaf in sunlight to decompose 0·11 c.c. of carbonic acid ($C\,O_2$) in an hour, find the volume of oxygen liberated in 5 hours by leaves whose total surface is 2500 sq. cm. Find also the quantity of carbon assimilated. [29.]

Hydrogen.

1. Calculate separately the quantities of zinc and sulphuric acid needed to prepare the following quantities of hydrogen, and the amount of zinc sulphate formed :—

(1) 6 grams. (2) 5 grams. (3) 1·6 oz. (4) 3·6 grams. (5) 4·2 oz. [6.]

2. How much zinc must be acted on by sulphuric acid to prepare the following quantities of hydrogen?

(1) 10 litres. (2) 15 litres. (3) 625 c.c. (4) 21 litres. (5) 3150 c.c. [6.]

3. How many grams of hydrogen would be produced by using the following amounts of zinc, and how much sulphuric acid would be necessary?

(1) 45·5 grams. (2) 71·5 grams. (3) 364 grams. (4) 227·5 grams. (5) 318·5 grams. [6.]

4. Find the quantities of iron and hydrogen chloride that will yield the following quantities of hydrogen :—

(1) 8 grams. (2) 1·8 oz. (3) 25 litres. (4) 5·6 grams. (5) 3125 c.c. [8.]

5. What would be the difference in volume between the gas obtained by acting on (1) 25·48 grams of iron, (2) the same amount of zinc, with sulphuric acid? [6, 7.]

6. How much sulphur is there—

(1) in 805 grams of zinc sulphate ($Zn\,SO_4$)?
(2) in 147 grams of sulphuric acid ($H_2\,SO_4$)?

7. The density of hydrogen referred to air is 0·0693; find the density of air, taking that of hydrogen as the unit.

8. How much zinc must be dissolved in hydrogen sulphate to liberate 60 litres of hydrogen at 15° C.? [6.]

9. How much zinc is required to prepare 266 litres of hydrogen at 7° C. and 750 m.m. pressure? [6.]

10. How much iron must be acted on by sulphuric acid to

furnish 55 litres of hydrogen at 13° C. and 570 m.m. pressure? [7.]

Nitric Acid.

1. Find the amounts of nitre and sulphuric acid required to prepare the following quantities of hydrogen nitrate :—

(1) 9 grams. (2) 56 grams. (3) 21 oz. (4) 15·75 grams. (5) 12·6 grams. [12.]

2. What mass of nitric acid would be obtained by using the following quantities of potassium nitrate, and how much hydrogen potassium sulphate would be formed?

(1) 15·15 grams. (2) 111·1 grams. (3) 70·7 oz. (4) 212·1 grams. (5) 282·8 grams. [12.]

3. Calculate the quantities of nitric anhydride and water in 49 grams of pure hydrogen nitrate ($H\,N\,O_3$).

4. Nitric acid of density 1·5 contains 79·7 per cent. of nitric anhydride; what percentage of $H\,N\,O_3$ does it contain?

5. How much hydrogen nitrate is required to form a neutral solution of nitre with 24 grams of caustic potash? [66.]

6. Find the quantity of metallic nitrate formed by dissolving in nitric acid—

(1) 13 grams of zinc.
(2) 58 grams of silver oxide.
(3) 267 grams of lead carbonate.
(4) 35 grams of bismuth.

7. How much ammonia must be neutralized by nitric acid to form 400 grams of ammonium nitrate? [13.]

8. How many grams of nitric acid containing 67·2 per cent. of $H\,N\,O_3$ will neutralize 54·4 grams of ammonia containing 36 per cent. of $N\,H_3$? [13.]

Nitrogen Monoxide.

1. Find what mass of nitrogen monoxide can be obtained by heating the following masses of ammonium nitrate :—

(1) 60 grams. (2) 100 grams. (3) 26 oz. (4) 14 grams. (5) 22 grams. [14.]

2. What quantities of ammonium nitrate will furnish the following volumes of nitrogen monoxide?

(1) 10 litres. (2) 625 c.c. (3) 15 litres. (4) 3375 c.c. (5) 6875 c.c. [14.]

3. How much phosphorus pentoxide is formed by burning 347·2 grams of phosphorus in excess of nitrous oxide, and what volume of nitrogen is liberated? [15.]

4. In what volume of oxygen would 7 grams of sulphur have to be burnt in order to form as much sulphur dioxide (SO_2) as when the combustion takes place in 12 litres of nitrous oxide? [16, 17.]

5. How much potassium oxide is formed by burning in nitrogen monoxide potassium enough to liberate 3125 c.c. of nitrogen? [18.]

6. How much ammonium nitrate will yield 19 litres of nitrogen monoxide at 7° C. and 875 m.m.? [14.]

7. What volume of nitrogen monoxide at 26° C. and 874 m.m. will 564·48 grams of ammonium nitrate yield? [14.]

8. What volume of nitrogen, estimated at 0° C. and 760 m.m., would be obtained by the decomposition of 320 c.c. of nitrous oxide at 17° C. and 551 m.m. pressure?

9. One volume of water at zero dissolves 1·305 volumes of nitrous oxide, but only 0·608 volume at 24° C. What volume of nitrous oxide would be obtained by heating 10 litres of water, saturated with the gas at zero, to 24° C., and what volume would it occupy at 0° C. under the same pressure?

10. How many grams of gas would be lost by allowing a jar of nitrous oxide to stand over 3125 grams of water at 0° C. till the latter was saturated?

Nitrogen Dioxide.

1. Find what quantities of copper and nitric acid will furnish the following quantities of nitrogen dioxide, and the mass of copper nitrate formed in each case:—
(1) 12 grams. (2) 18 oz. (3) 7·2 grams. (4) 13·2 grams.
(5) 10·8 grams. [19.]

2. How much copper must be acted on by nitric acid to prepare the following quantities of nitrogen dioxide?
(1) 10 litres. (2) 6250 c.c. (3) 35 litres. (4) 4375 c.c.
(5) 13¾ litres. [19.]

3. What mass of nitric oxide is theoretically just sufficient for the combustion of 93 grams of phosphorus? What volume of nitrogen would remain? [21.]

4. What volume of nitric oxide at 39° C. and 800 m.m. will

be obtained by acting on 74·676 grams of copper with hydrogen nitrate? [19.]

5. How much zinc must be dissolved in nitric acid to prepare 23 litres of nitric oxide at 26° C. and 570 m.m.? [20.]

6. How many grams of copper must be dissolved in nitric acid to prepare 11 litres of nitrogen dioxide at 13° C. and 950 m.m.? [19.]

7. Compare the masses of copper nitrate and zinc nitrate formed in preparing a litre of nitric oxide by acting on these metals respectively with nitric acid. [19, 20.]

8. Compare the masses of nitrogen in equal masses of nitrogen dioxide and monoxide.

9. Find the mass of the nitrogen trioxide obtained by cooling a mixture of 25 litres of nitric oxide with 6250 c.c. of oxygen. [22.]

10. What volume of oxygen must be added to nitric oxide to form as much nitrogen tetroxide as is obtained by heating 92·68 grams of lead nitrate? [23, 24.]

Ammonia.

1. Find the quantities of lime and sal-ammoniac necessary to prepare the following quantities of ammonia, and the mass of calcium chloride formed:—

(1) 17 grams. (2) 5·1 oz. (3) 42·5 grams. (4) 18·7 oz. (5) 35·7 grams. [25.]

2. What volumes of ammonia will be obtained by heating the following quantities of lime with a corresponding amount of sal-ammoniac?

(1) 1·568 gram. (2) 12·544 grams. (3) 7·84 grams. (4) 10·976 grams. (5) 68·992 grams. [25.]

3. How much water is formed by the combustion of 5 litres of ammonia in oxygen? [67.]

4. How many c.c. of hydrogen are set free by the decomposition of 4·76 grams of ammonia gas with the electric spark?

5. Find the pressure of the mixed gases liberated from a litre of ammonia gas at 760 m.m., assuming the temperature and volume to be the same as before decomposition.

6. How much lime must be heated with sal-ammoniac to prepare 95 litres of ammonia at 21° C. and 750 m.m.? [25.]

7. Find what volume of ammonia at 48° C. and 874 m.m. is obtained by heating 586·04 grams of ammonium chloride with quicklime. [25.]

8. How much zinc must be dissolved in nitric acid to liberate 8·5 grams of ammonia? [26.]

9. Water at 0° C. dissolves 1149 times its volume of ammonia gas; what mass of ammonia at 950 m.m. would 625 grams of water at 0° C. dissolve?

10. The specific gravity of a solution of ammonia containing 20 per cent. of gas is 0·925 at 14° C., and that of a solution containing 30 per cent. of gas is 0·897. Find the percentage of gas in a solution whose specific gravity is 0·911.

Carbon Dioxide.

1. How much marble and hydrogen chloride will suffice to prepare the following quantities of carbon dioxide, and how much calcium chloride will be formed?

(1) 33 grams. (2) 15·4 grams. (3) 26·4 grams. (4) 61·6 grams. (5) 96·8 grams. [28.]

2. Find what volume of carbon dioxide will be furnished by the following quantities of calcium carbonate:—

(1) 17·92 grams. (2) 3·92 grams. (3) 26·88 grams. (4) 12·544 grams. (5) 67·2 grams. [28.]

3. How much carbon must be burnt in oxygen to form—

(1) 25 litres of carbon dioxide at 0° C. and 760 m.m.
(2) 100 „ „ „ 21° C. and 760 m.m.? [29.]

4. Find the mass of the carbon deposited from carbon dioxide on burning in it 6 grams of magnesium wire. [68.]

5. What mass of carbon dioxide will precipitate 75 grams of calcium carbonate from lime-water, assuming no soluble carbonate to be formed? [30.]

6. Find the mass of carbon dioxide formed by the combustion of 150 cwt. of anthracite containing 94 per cent. of carbon.

7. How much marble will furnish 38 litres of carbon dioxide at 13° C. and 825 m.m.? [28.]

8. What volume of carbon dioxide at 14° C. and 820 m.m. can be obtained from 87·36 grams of marble? [28.]

9. Water at 15° C. dissolves its own volume of carbon dioxide, whatever the pressure; what mass of the gas can be

dissolved under a pressure of 3 atmospheres in 25 litres of water at 15° C.?

10. It is said to be injurious to breathe air containing 0·07 per cent. of carbon dioxide by volume. If a man exhale 260 grams of carbon dioxide in 8 hours, and remain for 8 hours in an unventilated room having a capacity of 23,000 litres, will the air in the room at the end of that time be fit to breathe or not?

Carbon Monoxide.

1. What mass of carbon monoxide is obtainable from the following amounts of oxalic acid?

(1) 9·9 grams. (2) 22·5 oz. (3) 13·5 grams. (4) 31·5 grams. (5) 24·75 grams. [33.]

2. What volume of carbon monoxide at 0° C. and 760 m.m. can be obtained from the following quantities of oxalic acid?

(1) 80·64 grams. (2) 141·12 grams. (3) 5·544 grams. (4) 27·72 grams. (5) 3·528 grams. [33.]

3. How much caustic soda is needed to purify the carbon monoxide obtainable from 63 grams of oxalic acid from carbon dioxide? [31, 33.]

4. How much carbon is necessary to reduce 55 grams of carbon dioxide to carbon monoxide? [34.]

5. The gases obtained by heating oxalic acid with strong sulphuric acid in a flask are passed through a solution of potassium hydrate; find the gain in mass of the solution when the flask has lost 54 grams. [32, 33.]

6. How much formic acid will yield as much carbon monoxide as can be got from 135 grams of oxalic acid? [33, 36.]

7. Find the volume at 0° C. and 855 m.m. of the carbon monoxide yielded by 90·72 grams of oxalic acid. [33.]

8. How much oxalic acid must be heated with hydrogen sulphate to prepare 145 litres of carbon monoxide at 17° C. and 950 m.m.? [33.]

9. A mixture of carbon monoxide and dioxide is exploded with 5 c.c. of oxygen, and the whole of the 15 c.c. which remain is absorbed by caustic potash. Find the composition of the mixture. [35.]

10. Find the volume at 0° C. and 760 m.m. of the gaseous mixture obtained from 100·8 grams of oxalic acid by heating it with hydrogen sulphate. [33.]

Marsh Gas.

1. What quantities of sodium acetate and caustic soda will furnish the following quantities of marsh gas, and how much sodium carbonate will be formed?

(1) 8 grams. (2) 2·8 grams. (3) 4·4 grams. (4) 5·2 grams. (5) 14·4 grams. [37.]

2. Find the volume of marsh gas obtained by heating the following masses of sodium acetate with caustic soda:—

(1) 73·472 grams. (2) 2·296 grams. (3) 128·576 grams. (4) 55·104 grams. [37.]

3. How much sodium acetate is needed to prepare 190 litres of marsh gas at 14° C. and 750 m.m.? [37.]

4. What volume of marsh gas at 0° C. and 779 m.m. could be obtained by using 941·36 grams of sodium acetate? [37.]

5. Calculate the percentage of hydrogen in marsh gas.

6. How much carbon is there in 1000 litres of marsh gas at 63° C. and 760 m.m. pressure?

7. 20 grams of marsh gas are burnt in air; what are the products of combustion? [38.]

8. Find the volume of marsh gas in 280 cubic feet of coal gas containing 41·5 per cent. of it by volume.

9. What volume of carbon dioxide would remain after exploding 7 c.c. of marsh gas with 14 c.c. of oxygen? [38.]

10. A mine contains 104,000 cubic feet of marsh gas and 1,000,000 cubic feet of atmospheric air at 0° C. and 760 m.m. pressure. What will be the volume of the gases at the same temperature and pressure after the explosion of the mixture? [38.]

Olefiant Gas.

1. How much alcohol is needed to prepare the following quantities of olefiant gas?

(1) 42 grams. (2) 49 grams. (3) 16·8 oz. (4) 15·4 grams. (5) 91 grams. [39.]

2. Calculate the mass of water abstracted from alcohol in forming the following volumes of olefiant gas:—

(1) 10 litres. (2) 2750 c.c. (3) 55 litres. (4) 65 litres. [39.]

3. What are the products of combustion of 21 grams of olefiant gas (*a*) in air, (*b*) in chlorine? [40, 41.]

4. Find the composition of the gas remaining after the explosion of 5 c.c. of olefiant gas with 20 c.c. of oxygen. [40.]

5. Compare the quantities of water formed by the combustion in oxygen of (*a*) equal volumes, (*b*) equal masses of marsh gas and olefiant gas. [38, 40.]

6. Find the mass of carbon in 19 litres of olefiant gas at 0° C. and 800 m.m.

7. How much alcohol is required to prepare 19 litres of olefiant gas at 26° C. and 875 m.m.? [39.]

8. Find what volume of olefiant gas at 14° C. and 779 m.m. can be obtained by decomposing 75·348 grams of alcohol with hydrogen sulphate. [39.]

9. What mass of olefiant gas would be obtained by the union of 312½ litres of hydrogen with an equal volume of acetylene? [42.]

10. A mixture of marsh gas and olefiant gas is exploded with 9 c.c. of oxygen; on treating the 6 c.c. of gas which remain with caustic potash, their volume is reduced to 2 c.c. Find the composition of the mixture. [38, 40.]

Chlorine.

1. Find the quantities of sodium chloride, hydrogen sulphate, and manganese dioxide needed to prepare the following quantities of chlorine:—

(1) 28·4 grams. (2) 42·6 grams. (3) 63·9 oz. (4) 78·1 grams. (5) 99·4 grams. (6) 125 litres. (7) 6250 c.c. (8) 15 litres. [43.]

2. What masses of sodium sulphate and manganese sulphate are produced in preparing the following quantities of chlorine?

(1) 56·8 grams. (2) 4375 c.c. (3) 92·3 grams. [43.]

3. If one volume of water at 15° C. dissolve 2·37 volumes of chlorine, find the mass of gas contained in 500 litres of water saturated at this temperature.

4. How much metallic chloride is formed by the direct combination with chlorine of—

(1) 100 grams of mercury. (2) 127 grams of copper. (3) 61 grams of antimony?

5. If 123·5 grams of chlorine at 0° C. occupy the same space as 50 grams of air, calculate the density of chlorine relatively to air.

6. What mass of chlorine is formed by passing 29·2 grams of gaseous hydrogen chloride over excess of heated manganese dioxide? [44.]

7. Find the mass of sodium chloride needed to prepare 25 litres of chlorine at 13° C. and 836 m.m. [43.]

8. What volume of chlorine at 0° C. and 750 m.m. is obtained by heating 98·28 grams of sodium chloride with manganese dioxide and hydrogen sulphate? [43.]

9. Find how much iodine is liberated by adding 60 grams of chlorine water containing 0·71 per cent. of chlorine by mass to excess of solution of potassium iodide. [45.]

10. What mass of chlorine would replace 32 grams of bromine in a compound of the latter?

Hydrogen Chloride.

1. Find the masses of sodium chloride and hydrogen sulphate needed to prepare the following quantities of hydrogen chloride, and the mass of hydrogen sodium sulphate formed:—

(1) 73 grams. (2) 6·57 oz. (3) 15·33 grams. (4) 80·3 grams. (5) 109·5 grams. [46.]

2. How much sodium chloride and hydrogen sulphate will serve to prepare the following volumes of gaseous hydrogen chloride?

(1) 10 litres. (2) 625 c.c. (3) 125 litres. (4) 15$\frac{5}{8}$ litres. [46.]

3. What mass of hydrogen chloride is formed by the combination of 62$\frac{1}{2}$ litres of hydrogen with an equal volume of chlorine? [48.]

4. Find the mass of the chloride formed by acting with hydrochloric acid on—

(1) 59 grams of tin. (2) 112 grams of caustic potash. (3) 159 grams of copper oxide. (4) 34 grams of ammonia.

5. How much platinum chloride is obtained by dissolving 3·94 grams of platinum in aqua regia?

6. Compare the masses of hydrogen sodium sulphate and

hydrogen potassium sulphate formed in preparing equal masses of hydrogen chloride with sodium chloride and potassium chloride. [46, 47.]

7. What volume of gaseous hydrogen chloride at 39° C. and 836 m.m. would be got by acting on 1891·89 grams of sodium chloride with hydrogen sulphate? [46.]

8. How much potassium chloride is needed for the preparation of 38 litres of gaseous hydrogen chloride at 25° C. and 750 m.m.? [47.]

9. The percentage of hydrogen chloride in hydrochloric acid of density 1·21 being 42·43, and in acid of density 1·13 being 26·26; find the percentage of hydrogen chloride in acid of density 1·15.

10. How much solution of hydrogen chloride, containing 30 per cent. of the gas, will neutralize 102 grams of solution of ammonia containing 35 per cent. of ammonia gas? [49.]

Sulphur.

1. Find the masses of copper and hydrogen sulphate needed to prepare the following quantities of sulphur dioxide, and the mass of copper sulphate formed:—

(1) 16 grams. (2) 17·92 grams. (3) 25·6 grams. (4) 70·4 grams. (5) 76·8 grams. [50.]

2. What volume of sulphur dioxide would be obtained by acting with strong sulphuric acid on the following quantities of copper?

(1) 14·224 grams. (2) 17·78 grams. (3) 24·892 grams. (4) 99·568 grams. (5) 170·688 grams. [50.]

3. Calculate the mass of sulphur dioxide in 123 grams of hydrogen sulphite. [55.]

4. How much hydrogen sulphate would be obtained from 96 grams of sulphur dioxide by treatment with water and nitrogen tetroxide? [52.]

5. How much sulphur must be burnt in air to form as much sulphur dioxide as is obtained by acting on 25 grams of mercury with hydrogen sulphate? [51, 16.]

6. How much mercury and sulphuric acid are required to prepare 9500 c.c. of sulphur dioxide at 21° C. and 875 m.m.? [51.]

Hydrogen Sulphide.

1. Find the masses of iron sulphide and hydrogen sulphate needed to prepare the following masses of hydrogen sulphide, and the quantities of iron protosulphate formed:—

(1) 10·2 oz. (2) 11·22 grams. (3) 37·4 grams. (4) 71·4 grams. (5) 132·6 grams. [53.]

2. Find the masses of iron sulphide and sulphuric acid needed to prepare the following volumes of hydrogen sulphide.

(1) 5 litres. (2) 2 litres. (3) 6250 c.c. [53.]

3. Calculate the mass of sulphide precipitated by adding 17 grams of hydrogen sulphide (in solution) to excess of acidulated solution of—

(1) Copper sulphate. (2) Mercuric chloride. (3) Bismuth nitrate.

4. What mass of hydrogen sulphide will precipitate as much copper sulphide from a solution of a copper salt as would be formed by heating 127 grams of copper with excess of sulphur? [54.]

5. What are the products of combustion in oxygen or air of 51 grams of hydrogen sulphide? [55.]

6. How much sulphur is deposited on decomposing 5 litres of hydrogen sulphide with red-hot platinum?

7. Assuming combination to take place, what mass of hydrogen sulphite would be formed by exploding 9 litres of oxygen with 6 litres of hydrogen sulphide? [55.]

8. Calculate the mass of iron sulphide decomposed to prepare 25 litres of hydrogen sulphide at 13° C. and 798 m.m. [53.]

9. What volume of hydrogen sulphide at 7° C. and 750 m.m. will be obtained by acting on 216·216 grams of iron sulphide with sulphuric acid? [53.]

10. One volume of water at 15° C. dissolves about 3·2 volumes of hydrogen sulphide; how much iron sulphide will be needed to prepare gas enough to saturate 375 litres of water at this temperature, and under 760 m.m. pressure, assuming no loss to occur?

Phosphorus.

1. If one pound of bone-ash contain 85 per cent. of tricalcic

phosphate ($Ca_3\ 2\,P\,O_4$), how much phosphorus could be obtained from 100 pounds of bone-ash?

2. What mass of phosphine is obtained by acting on 62 grams of phosphorus with caustic potash? [56.]

3. Find what mass of calcium hypophosphite is formed by heating 6·2 grams of phosphorus with excess of lime-water. [57.]

4. How much phosphorus is contained in—

(1) 27·4 grams of microcosmic salt ($Na\,H_4N\,H\,P\,O_4$).
(2) 51 grams of sodium metaphosphate ($Na\,P\,O_3$).
(3) 33·25 grams of sodium pyrophosphate ($Na_4P_2O_7$)?

5. How much sodium metaphosphate would be formed by heating 41·1 grams of microcosmic salt? [58.]

6. Find what quantity of sodium phosphate must be heated to furnish 93·1 grams of sodium pyrophosphate. [59.]

7. Compare the percentages of phosphorus in hydrogen phosphite ($H_3P\,O_3$), and hydrogen phosphate ($H_3P\,O_4$).

8. What volume of oxygen is needed to produce 68·6 grams of hydrogen phosphate by combustion of phosphine? [64.]

9. What volume of phosphine is obtained by decomposing 22·96 grams of hydrogen phosphite by heat? [70.]

10. How much phosphorus oxychloride is produced by decomposing 83·4 grams of the pentachloride with water? [71.]

CHAPTER X.

WATER.

Connection of boiling-point with pressure.

Water boils when the tension of its vapour is equal to the pressure on its surface. Hence, the less the pressure, the lower the boiling-point will be.

The temperature at which water boils under a given pressure, or the pressure under which water will boil at a given temperature, is found from a table of the tensions of aqueous vapour at various temperatures.

Ex. *Water boils at* 40° *C. under a pressure of* 54·906 *m.m., and at* 50° *C. under a pressure of* 91·982 *m.m.; find the pressure under which water will boil at* 45° *C.*

An increase of 37·076 m.m. in the pressure raises the boiling-point from 40° C. to 50° C., *i.e.* 10° C. We wish to know what increase of pressure will raise the boiling-point 5° C. This is given by the proportion—

$$10 : 5 :: 37{\cdot}076 : x,$$

$$\text{whence } x = \frac{5 \times 37{\cdot}076}{10}$$

$$= 18{\cdot}538.$$

Hence the pressure required is—

$$54{\cdot}906 \text{ m.m.} + 18{\cdot}538 \text{ m.m.}$$

$$\text{or } 73{\cdot}444 \text{ m.m.}$$

[Reference to a table shows that the pressure at 45° C. is really 71·391 m.m. The above result is, therefore, too large by 2·05 m.m.; this is because we have assumed the change of pressure to be proportional to the 10° change of temperature, whereas it is strictly so only for *small* changes of temperature.]

EXAMPLES X.

1. How many grams of hydrogen and oxygen would be furnished by the electrolysis of 45 grams of water? [27.]

2. A certain mass of water yields on decomposition 2·5 grams of hydrogen; what mass of oxygen would it yield?

3. What mass of oxygen will just combine with 3 grams of hydrogen to form water?

4. How many c.c. of hydrogen at 0° C. will just combine with 10 c.c. of oxygen at 0° C. to form water?

5. What volume of oxygen at 0° C. will just suffice for the complete combustion of 192 c.c. of hydrogen at 15° C.?

6. How much water must be decomposed to furnish 5·2 grams of hydrogen, and what mass of oxygen is liberated?

7. How much water would furnish 56 grams of oxygen, and what mass of hydrogen would be given off?

8. How many litres of hydrogen would be furnished by 24·57 grams of water at 7° C., the gas keeping that temperature?

9. How much sodium would liberate 5 grams of hydrogen from water, and how much sodium oxide would be formed. [9.]

10. Compare the masses of caustic potash and caustic soda formed by decomposing equal masses of water with potassium and sodium respectively. [9, 10.]

11. If 93 grams of sodium oxide are produced, what mass of hydrogen is liberated, and how much water decomposed? [9.]

12. Calculate separately the masses of hydrogen and magnetic oxide formed by oxidizing the following quantities of red-hot iron with steam:—

(1) 42 oz. (2) 54·6 grams. (3) 29·4 lbs. (4) 176·4 grams. [11.]

13. In the synthesis of water by weight, the bulb through which the hydrogen is passed is found finally to contain 12·7 grams of copper, and 4·1 grams of unreduced copper oxide.

How much oxide did the bulb originally contain, and how much water was formed? [4, 27.]

14. The vapour tension of water at 90° C. is 525·45 m.m., and 760 m.m. at 100° C.; at what temperature would water boil when the barometer indicates 713·09 m.m. pressure?

15. Water boils at 30° C. under a pressure of 31·548 m.m., and at 40° C. under a pressure of 54·906 m.m.; calculate the vapour tension of water at 35° C. (See note, p. 52.)

16. What temperature is indicated by a Centigrade thermometer surrounded by steam, when the atmospheric pressure is 736·545 m.m.? (See Ex. 14.)

17. In an atmosphere of uniform temperature 0° C., if the pressure fall 1 per cent. for a rise of 80 metres, calculate the height of a mountain at the top of which water boils at 92°·4 C.

18. Under the same conditions, at what temperature would water boil on a mountain 469·1 metres high?

CHAPTER XI.

NITROGEN AND AIR.

Hygrometric Formula.

The mass of a cubic metre of dry air at 0° C. and 760 m.m. is 1293 grams.

Now the density of aqueous vapour referred to air is 0·625, or $\frac{5}{8}$. Hence, the mass of a cubic metre of aqueous vapour, if it could exist at standard pressure and temperature, would be—

$$\frac{1293 \times 5}{8} \text{ grams.}$$

Hence, the mass of V cubic metres of aqueous vapour at t° C. and pressure p m.m. (the vapour tension for that temperature), in other words the quantity of water in V cubic metres of air saturated with moisture at t° C., is—

$$\frac{1293 \times p\mathrm{V} \times 273}{760 \times (273 + t)} \times \frac{5}{8} \text{ grams. (Chap. V.)}$$

p is found from a table of tensions of aqueous vapour.

Ex. (1). *Find the amount of aqueous vapour in a cubic metre of air saturated with moisture at* 5° *C.*

The table gives for 5° C., $p = 6{\cdot}534$ m.m.

$$\therefore \text{ mass required} = \frac{1293 \times 6{\cdot}534 \times 273}{760 \times 278} \times \frac{5}{8} \text{ grams.}$$
$$= 6{\cdot}825 \text{ grams, nearly.}$$

Ex. (2). 182 *c.c. of air saturated with moisture are at* 0° *C. and* 760 *m.m. pressure. What will be the volume of this air, still saturated with moisture, at* 5° *C., and under the same pressure?*

At 0° C. the tension of aqueous vapour is 4·6 m.m.;

hence the pressure of the air alone is (760 — 4·6) m.m., *i.e.* 755·4 m.m. (p. 16, § 3).

Similarly the pressure of the air alone at 5° C. is (760 — 6·534) m.m., or 753·466 m.m.

If, then, v be the volume required,

$$\frac{753\cdot466\, v}{273+5} = \frac{182 \times 755\cdot4}{273}$$

$$\therefore v = \frac{182 \times 755\cdot4 \times 278}{753\cdot466 \times 273} \text{ c.c.}$$

$$= 185\cdot8 \text{ c.c. approximately.}$$

[Throughout the following examples nitrogen is assumed to form $\frac{4}{5}$ of the atmosphere by volume].

EXAMPLES XI.

1. What volume of oxygen is there in 15 litres of air?

2. Calculate the masses of nitrogen and oxygen in 250 litres of air (*a*) at 0° C. (*b*) at 7° C., the pressure being 760 m.m.

3. Find the masses of nitrogen and oxygen in 31·25 litres of air at 0° C., when the barometer stands at 722 m.m.

4. On passing air over red-hot copper 35·616 grams of copper oxide are formed; what volume of nitrogen is left? [4.]

5. What volume of air at standard pressure and temperature will furnish oxygen enough to oxidize 45·5168 grams of copper? [4.]

6. How much copper oxide is formed during the decomposition of 6250 c.c. of air by passing it over red-hot copper? [4.]

7. What volume of nitrogen would remain if 69·44 grams of phosphorus were completely burnt in a quantity of air theoretically just sufficient?

8. How much phosphorus pentoxide would be formed in decomposing 125 litres of air with phosphorus?

9. 125 litres of air are drawn through tubes containing

pumice moistened with sulphuric acid by means of an aspirator. The tubes weigh before the experiment 100 grams, and 100·5 grams at its conclusion. Calculate the amount of aqueous vapour in a cubic metre of air at the time.

10. A cubic metre of air saturated with moisture at 25° C. contains 22·5 grams of water, but only 5·4 grams at 0° C. How much rain would be deposited from 5000 litres of air on thus cooling?

11. Calculate approximately the mass of aqueous vapour in a cubic metre of air saturated with moisture—

(1) At 0° C., the tension of aqueous vapour being 4·6 m.m.
(2) At 11° C., „ „ „ 9·792 mm.
(3) At 40° C., „ „ „ 54·906 m.m.
(4) At −6° C., „ „ „ 2·876 m.m.

12. Air at 12° C. containing $\frac{4}{5}$ of the maximum quantity of aqueous vapour for that temperature is cooled to zero. How much rain will be deposited from 100 cubic metres of it? (Tension of aqueous vapour at 12° C., 10·457 m.m.)

13. A cubic metre of air at 15° C. is passed through a tube containing pumice moistened with sulphuric acid. After the passage of the air the tube weighs 3·95 grams more than it did before. Taking the tension of aqueous vapour to be 1·27 cm., what decimal fraction of the maximum quantity of moisture possible at its temperature did the air contain?

14. 364 c.c. of air saturated with moisture are at 0° C. and 760 m.m.; what will be the volume of this air, still saturated with moisture, at 26° C. and 830 m.m.?

Tension of aqueous vapour at 26° C., 2·5 cm.
„ „ „ 0°.C., 0·46 cm.

15. Find approximately the change of mass of 1000 cubic inches of air when the atmospheric pressure rises from 29·55 inches to 30·33 inches, the temperature being constantly 0° C. Assume 100 cubic inches of air at the former pressure to weigh 31 grains.

CHAPTER XII.

THE CRITH.

To avoid the labour of reducing the mass of a given quantity of gas to grams, it has been found convenient to introduce a new unit, called the *crith*. It is the mass of a litre of hydrogen at 0° C. and 760 m.m., *i.e.* 0·0896 gram.

It is easily seen that the number of criths in a litre of any other gas at 0° C. and 760 m.m. is equal to the number expressing the density of that gas referred to hydrogen. Thus a litre of nitrogen at standard pressure and temperature weighs 14 criths, a litre of carbon dioxide 22 criths, and so on.

To find the number of criths in a litre of gas at any other pressure and temperature we must first find its volume at 0° C. and 760 m.m. and then multiply by its normal density.

Ex. (1). *Express in criths the mass of 7 litres of oxygen at 7° C. and* 950 *m.m. pressure.*

The volume, reduced to standard pressure and temperature, is—

$$v = \frac{7 \times 273 \times 950}{(273+7) \times 760} \text{ litres}$$
$$= \frac{273}{32} \text{ litres.}$$

Hence, multiplying by 16, the density of oxygen at 0° C. and 760 m.m., we have

$$\text{mass required} = \frac{273}{32} \times 16$$
$$= 136{\cdot}5 \text{ criths.}$$

Ex. (2). *What mass of sulphur is contained in* 51 *criths of hydrogen sulphide, and what volume does this quantity of gas occupy at standard pressure and temperature?*

We know from the formula of the gas that 34 parts by weight of H_2S contain 32 parts by weight of sulphur.

$$\therefore\ 1 \text{ crith of } H_2S \text{ contains } \frac{32}{34} \text{ criths of sulphur.}$$

$$\therefore\ 51 \text{ criths } \quad ,, \quad \text{contain } \frac{32}{34} \times 51 \quad ,,$$

or 48 criths.

Again, 1 litre of H_2S weighs 17 criths.

∴ 51 criths of H_2S occupy 3 litres.

This, then, is the volume required.

EXAMPLES XII.

1. Express in criths the mass at 0° C. and 760 m.m. of—

(1) 2 litres of ammonia. (2) 500 c.c. of nitrogen monoxide. (3) 4 litres of chlorine. (4) 800 c.c. of nitrogen dioxide.

2. Express in criths the mass of—

(1) 20 litres of nitrogen at 7° C. and 760 m.m.
(2) 19 litres of hydrogen at 0° C. and 800 m.m.
(3) 950 c.c. of carbon dioxide at 13° C. and 800 m.m.

3. How many criths of water are formed by the combination of a litre of hydrogen with half the volume of oxygen? [27.]

4. How many criths of nitrogen remain after removing the oxygen from 45 criths of nitrogen dioxide?

5. What volume of carbon dioxide ($C\,O_2$) would deposit 36 criths of carbon on burning in it magnesium wire? [68.]

6. What volume of sulphur dioxide ($S\,O_2$) is formed by burning 11 criths of sulphur in oxygen?

7. How many grams of carbon monoxide ($C\,O$) contain 75 criths of carbon?

8. How many criths of nitrogen are contained in 54·4 grams of gaseous ammonia ($N\,H_3$)?

CHAPTER XIII.

HEAT CONSIDERED AS A FORM OF ENERGY.

1. *Force.*

Any cause which tends to alter a body's state of rest or of uniform motion in a straight line is called force. We are familiar with it in the effort needed to support a weight, to keep a spring extended, to stop a cricket-ball, &c.; or in the passive resistance by which a table prevents bodies upon it from falling to the ground.

2. *Work.*

When force is used to overcome another force, work is done. The work done is measured by the product of the force overcome and the distance through which its point of application is moved in the direction of the force. Or

$$\text{work} = \text{force} \times \text{space traversed}.$$

If the work done in lifting a pound through one foot against its own weight be called a foot-pound, six foot-pounds of work will have to be expended to lift a pound through six feet, or six pounds through one foot, or two pounds through three feet. Here the force overcome is the weight of a pound; since this varies from one part of the earth to another the foot-pound is an indefinite unit unless the latitude and height above the sea-level be stated. An *erg* is the work which is done when the point of application of a dyne (see p. 11) is moved through one centimetre. Since this is independent of the locality, it is called an absolute unit.

Work is also done when force is used to set a mass in motion, *e.g.* when a gun is fired. The work then done is measured by the product of the mass of the body moved and the square of the velocity imparted to it. It is usual to choose the units of length, mass, and time, so that—

$$\text{work} = \tfrac{1}{2}\,\text{mass} \times (\text{velocity})^2.$$

3. *Energy.*

Energy is capacity of doing work. It may be due to various causes; a cannon-ball in motion can do work in virtue of its mass and its motion; gunpowder can do work in virtue of the chemical affinity of the substances of which it is composed; a weight raised above the ground can do work as it descends, in virtue of its position. Energy due to motion is called *kinetic energy;* that due to position or molecular arrangement is called *potential energy.* The kinetic energy of a body is

$$\tfrac{1}{2}\,\text{mass} \times (\text{velocity})^2.$$

4. *Conservation of Energy.*

The kinetic energy possessed by a body may be transformed into potential energy, or *vice versâ,* but so long as no work is done by or upon the body the total amount of energy it possesses will remain the same. Consider, for instance, the case of 1 cwt. raised 10 feet above the earth's surface. The work expended in lifting it is 1120 foot-pounds, and if allowed to fall it could evidently do this amount of work in raising another weight, *i.e.* it possesses 1120 foot-pounds of potential energy. If it fall without doing work, it will lose potential energy in consequence of its loss of height, but this potential energy will be replaced by kinetic energy, which increases as the velocity increases. And since the force moving it is its weight, and this force acts through 10 feet, the work done by gravity in making the body fall will clearly be the same as that done in lifting it against gravity, when its weight was overcome through 10 feet. Thus on reaching the ground it will be found to possess 1120 foot-pounds of kinetic, but no potential energy.

This is an instance of the law known as the Conservation of Energy; it asserts that the total amount of energy in any system of bodies is always the same, if no work is done by the system against external forces nor expended on the system by external forces.

5. *Mechanical Equivalent of Heat.*

When the cwt. considered in the last paragraph strikes the ground, its motion is arrested and its kinetic energy disappears. Clearly, also, it can have no potential energy due to position above the earth's surface. Experiment, however, shows that the law of conservation still holds good. A small portion of the body's energy will be accounted for by the work required to produce the sound of the shock and the indentation of the earth, but the greater part takes the form of heat. Joule has found by experiment that 772 foot-pounds of energy in the form of heat will raise the temperature of a pound of water 1° F. Or, the mechanical equivalent of heat, taking the foot, pound, and degree Fahrenheit as units, is 772 foot-pounds. It follows that to raise a pound of water 1° C., *i.e.* $\frac{9}{5}$° F., about 1390 foot-pounds of heat are required. Lastly, if we take the metre, kilogram, and degree Centigrade as units, the mechanical equivalent of heat is about 427 kilogram-metres, *i.e.* 427 kilograms falling through 1 metre acquire kinetic energy enough to raise the temperature of a kilogram of water 1° C., if this energy be applied to the water in the form of heat.

6. *Latent Heat.*

We can now see more clearly what becomes of the heat absorbed or liberated during change of state, *e.g.* during the vaporization of water. Work has to be done in separating the molecules; and the heat-energy required for the purpose disappears as such, but is stored up in the steam as potential energy, reappearing as heat when condensation takes place.

7. *Ex.* (1). *A ball of lead weighing* 20 *lbs. falls from a height of* 139 *feet; how much heat is produced when it strikes the ground, assuming none of its energy to take other forms?*

Potential energy of the ball = 20 × 139 foot-pounds.
= 2780 foot-pounds.
∴ kinetic energy of ball on
reaching ground = 2780 foot-pounds.

But 1390 foot-pounds of energy will raise the temperature of 1 lb. of water 1° C.

∴ 2780 foot-pounds will raise 1 lb. of water 2° C.

Ex. (2). *Find the rise in temperature of the ball.*

Suppose it to be $x°$ C. Then the same amount of heat serves to raise 20 lbs. of lead (specific heat 0·03) through $x°$ C. as raises 1 lb. of water through 2° C.

$$\therefore 20 \times x \times 0{\cdot}03 = 2 \times 1.$$

$$\therefore x = \frac{2}{20 \times 0{\cdot}03}° \text{ C.}$$

$$= 3\tfrac{1}{3}° \text{ C.}$$

Or, without introducing an equation, heat sufficing to raise 1 lb. of water 2° C. will raise 1 lb. of lead $\frac{2}{0{\cdot}03}°$ C. and 20 lbs. of lead $\frac{2}{20 \times 0{\cdot}03}°$ C.

Ex. (3). *From what height must the ball fall in order to be melted by the shock, if its temperature be 6° C. at the moment of collision?*

Lead melts at 326° C.; hence the heat required to raise the temperature of the ball to melting-point is $20 \times 0{\cdot}03\ (326 - 6)$ units, or 192 units.

The latent heat of lead is about 5·4, hence the additional heat required to melt the ball is $20 \times 5{\cdot}4$ units, or 108 units.

Altogether, then, 300 units of heat are required, *i.e.* heat enough to raise the temperature of 300 lbs. of water 1° C. Therefore the height from which the ball must fall is 300×1390 feet, or 139,000 yards, about 79 miles.

[It need hardly be pointed out that the experimental data employed are insufficient for this case.]

8. *Heat of Combination.*

We have seen that the molecules of two substances capable of chemical combination possess potential energy in consequence of their affinity. After combination has occurred, they of course no longer possess potential energy; it has been transformed into heat, which shows itself in a rise of temperature when combination takes place. If the compound formed were decomposed again into the original substances, an amount of heat equal to that liberated would be reabsorbed, and the molecules would regain their potential energy. Some-times heat is absorbed during combination; in this case heat is liberated when the compound is decomposed.

In either case, the quantity of heat depends on the quantity of the substance formed or decomposed, and is different for different substances. Favre and Silbermann have determined these heat-changes in several cases; the following are some of their results, omitting decimals:—

I.

Substance.	Calories liberated when 1 gram combines with oxygen.
Hydrogen	34,462
Charcoal	8,080
Sulphur	2,260
Iron	1,353
Zinc	1,291
Copper	684
Lead	266
Silver	57

II.

Substance.	1 gram of anhydrous chloride formed.	Combination of 1 gram to form chloride in solution.
Sodium	4,126	4,101
Potassium	2,588	2,489
Iron	1,775	1,905
Zinc	1,547	1,740
Copper	923	1,078

The difference between the columns of Table II. is due to the heat absorbed on dissolving the anhydrous chloride in water. ("Ann. de Chem. et Phys.," 3rd Series, Vols. 34, 36, 37).

Ex. *How much heat is produced during the formation of 53 grams of copper oxide by burning the metal in oxygen?*

To form 53 grams of copper oxide we need a quantity of copper given by the proportion—

$$79{\cdot}5 : 53 :: 63{\cdot}5 : x.$$

$$\therefore x = \frac{53 \times 63{\cdot}5}{79{\cdot}5} \text{ grams.}$$

$$= \frac{127}{3} \text{ grams.}$$

$$\therefore \text{heat produced} = \frac{127}{3} \times 684 \text{ calories.}$$

$$= 28{,}956 \text{ calories.}$$

EXAMPLES XIII.[1]

1. A ball of iron weighing 5 lbs. produces by collision with the ground heat enough to raise 2 lbs. of water 1° F. How far did it fall?

2. Find the rise in temperature (Centigrade) of a ball of iron weighing 10 lbs. which strikes the ground after falling 382¼ ft.?

3. From what height must a copper ball weighing 5 lbs. fall that its temperature may be raised 1° F. by collision with the ground? (Sp. heat 0·095.)

4. An iron nail weighing 4 oz. is struck by a hammer weighing 22 lbs. with as much force as if the hammer had fallen freely from a height of 48¼ ft. Find the rise in temperature (Fahrenheit) of the nail.

5. How much heat is produced by the combustion of 6 grams of iron wire in oxygen?

6. Find the heat produced by the formation from elements of—

(1) 27 grams of water. (2) 5 grams of sulphur dioxide. (3) 4 grams of anhydrous copper chloride.

7. How much charcoal would have to be burnt in oxygen to furnish heat enough to raise the temperature of 404 grams of water 100° C.?

8. How many grams of water would be raised 1° C. by the heat produced in forming 1 gram of water?

9. Compare the masses of charcoal and hydrogen that must be burnt to produce a given quantity of heat.

10. Calculate the heat absorbed during solution in water of 1 gram of anhydrous copper chloride.

11. How much work is done in lifting—

(*a*) 3 kilograms of iron through a metre?
(*b*) 8 centigrams of iron through 14 metres?
(*c*) A ton of coal from the bottom of a pit 10 fathoms deep?

12. What is the kinetic energy of 4 kilograms moving at the rate of 45 kilometres an hour, and how much work must be expended to produce it?

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[1] In the following Examples, assume all the energy to be transformed into heat and to enter the body under consideration.

CHAPTER XIV.

MISCELLANEOUS EXAMPLES.

WE shall here work a few examples of a miscellaneous character in illustration of those to be found at the end of the chapter.

I. Calculate the percentage composition of quicklime (Ca O).

The formula shows that 56 parts by weight of quicklime contain 40 parts of calcium and 16 of oxygen. We have to find the quantities of these elements in 100 parts of quicklime.

They are evidently determined by the two proportions—

56 : 100 :: 40 : amount of calcium.
56 : 100 :: 16 : amount of oxygen.

$$\therefore \text{ percentage of calcium} = \frac{100 \times 40}{56} = 71{\cdot}43 \text{ q.p.}$$

$$\text{,, } \quad \text{oxygen} = \frac{100 \times 16}{56} = 28{\cdot}57 \text{ q.p.}$$

II. Find a formula for a substance with the following percentage composition,—

Silicon 20·28, *magnesium* 34·16, *oxygen* 45·56.

If we divide each of these numbers by the atomic weight of the element to which it is affixed, we shall clearly get numbers proportional to the number of atoms of each of these elements in a molecule of the compound.

$$\text{Now } \frac{20{\cdot}28}{28} = 0{\cdot}724\ldots, \frac{34{\cdot}16}{24} = 1{\cdot}42\dot{3}; \ \frac{45{\cdot}56}{16} = 2{\cdot}847\ldots$$

These numbers are not exact integral multiples one of another; but if we divide each of the others by the least, and take the nearest integer to the result as the true quotient, we get the most probable proportion of atoms of each element.

Thus $\frac{1{\cdot}42\dot{3}}{{\cdot}724} = 2$, $\frac{2{\cdot}847}{{\cdot}724} = 4$, approximately.

Hence the probable formula of the compound is $SiMg_2O_4$.

III. 5 c.c. of a mixture of marsh gas and olefiant gas are exploded with 14 c.c. of oxygen. There remain 9 c.c. of gas, of which 7 c.c. are absorbed by caustic potash; find the composition of the original mixture.

Olefiant gas requires three times its volume of oxygen for its combustion; marsh gas twice its volume. The carbon dioxide produced occupies twice the volume of the olefiant gas, and the same volume as the marsh gas (see equations 38 and 40).

Suppose the mixture to be formed of x c.c. of olefiant gas and y c.c. of marsh gas, so that $x + y = 5$. Then the carbon dioxide produced occupies $(2x + y)$ c.c.; whence $2x + y = 7$. This gives $x = 2, y = 3$.

And the volume of oxygen required to burn a mixture of 2 c.c. of olefiant gas and 3 c.c. of marsh gas is $(3 \times 2 + 2 \times 3)$ c.c., or 12 c.c. This accounts for the 2 c.c. of gas remaining after absorption by caustic potash—they were unused oxygen.

Or, without introducing algebra, we might have sought two numbers whose sum is 5, such that twice one of them together with the other amounts to 7.

MISCELLANEOUS EXAMPLES XIV.

1. Calculate the masses of the various constituent elements in 100 grams of each of the following substances :—

(1) Calcium carbonate. (2) Sulphur dioxide. (3) Magnesia. (4) Epsom salts. (5) Caustic soda. (6) Ammonium nitrate. (7) Ammonium carbonate. (8) Zinc carbonate. (9) Oxalic acid. (10) Hydrogen sodium sulphate.

2. Find the simplest formula of substances having the following percentage composition :—

(1) Sodium 32·79, aluminium 13·02, fluorine 54·19.
(2) Iron 34·4, arsenic 46·0, sulphur 19·6.
(3) Sodium 34·59, phosphorus 23·31, oxygen 42·1.

(4) Sodium 22·55, phosphorus 30·39, oxygen 47·06.
(5) Copper 34·6, iron 30·52, sulphur 34·88.
(6) Iron 25, oxygen 28·57, chromium 46·43.
(7) Sodium 22·77, boron 21·78, oxygen 55·45.
(8) Carbon 10·04, hydrogen 0·84, chlorine 89·12.
(9) Carbon 20, hydrogen 6·7, nitrogen 46·7, oxygen 26·6.
(10) Potassium 35·56, iron 17·02, carbon 21·89, nitrogen 25·53.

3. A certain compound weighs 25 grams, and contains 10 grams of calcium, 12 grams of oxygen, and 3 grams of carbon. Find a formula for it.

4. Find the empirical formula of a volatile liquid having the following percentage composition :—

Carbon 52·37, hydrogen 13·31, oxygen 34·61.

5. 0·318 gram of an organic acid yielded on analysis 0·8005 gram of carbonic acid, and 0·1422 gram of water. Find its empirical formula.

6. What is the water-equivalent of a copper calorimeter weighing 96 grams, the specific heat of copper being 0·095?

7. 6 kilograms of mercury at 100° C. are stirred with 2 kilograms of ice-cold water, and the final temperature is 9° C. Find the specific heat of mercury.

8. 100 grams of a mixture of iron and copper filings, raised to a temperature 100° C., are thrown into 300 grams of water at 0° C., contained in a silver vessel weighing 50 grams. The final temperature of the water is 3°·08 C.; determine the composition of the mixture. (Sp. heat of iron 0·114, copper 0·092, silver 0·056.)

9. What is the range Centigrade of a thermometer with a range of 19° F.?

10. Find the reading at 72° C. of a thermometer on which the graduations for the temperatures of melting ice and boiling water are marked 51 and 126 respectively.

11. The specific heat of iron being 0·1138, calculate its probable atomic weight from Dulong and Petit's law.

12. One gram of a certain metal when dissolved in dilute sulphuric acid liberates 200 c.c. of hydrogen gas; taking the volume of a gram of hydrogen as 11·2 litres, find the combining weight of the metal.

13. Calculate from the formula of common alum the proportions in which the salts composing it occur.

14. Find approximately the masses of water and anhydrous salt in 100 grams of crystals of the following composition:—

(1) $Ca\,Cl_2 + 6\,H_2O$. (4) $Mn\,Cl_2 + 4\,H_2O$.
(2) $Mg\,S\,O_4 + 7\,H_2O$. (5) $Fe\,Cl_2 + 4\,H_2O$.
(3) $Zn\,S\,O_4 + 7\,H_2O$.

15. Common soda crystals ($Na_2C\,O_3 + 10\,H_2O$) on exposure to the air alter their composition to $Na_2C\,O_3 + H_2O$. If 143 lbs. of soda-crystals undergo this change, what will be the mass of the resulting substance?

16. 280 grams of iron are placed in a solution containing 1000 grams of copper sulphate; how much copper sulphate will be left in solution when all the iron has been dissolved? [60.]

17. How much zinc will replace the lead in 10 grams of lead acetate in solution? [61.]

18. Calculate the mass of the copper oxide obtained by heating 50 grams of copper nitrate to redness. [69.]

19. How much corrosive sublimate must be stirred with 83 grams of potassium iodide, in order that the resulting mixture may consist solely of mercuric iodide and potassium chloride? [62.]

20. What mass of hydrogen chloride must be decomposed to yield enough chlorine to liberate 3·81 grams of iodine from potassium iodide? [45.]

21. Find the volume at 0° C. and 760 m.m. of each of the gases obtained by heating 67·2 grams of charcoal with sulphuric acid. [63.]

22. What volumes of the constituent gases would be obtained by decomposing—

(1) 24 c.c. of ammonia? (2) 50 c.c. of nitrogen dioxide? (3) 6 litres of nitrogen pentoxide? (4) 5 litres of hydrogen chloride? (5) 100 c.c. of nitrogen monoxide?

23. Find what volume of oxygen is required for the combustion of a litre of each of the following gases, and the volume of the gaseous product in each case:—

(1) **Carbon monoxide.** (2) **Olefiant gas.** (3) **Hydrogen sulphide.** (4) **Phosphuretted hydrogen.** [35, 40, 55, 64.]

24. How many c.c. of oxygen will be required to burn completely a mixture of 10 c.c. of marsh gas, and 20 c.c. of ethylene, and what volume of carbon dioxide will be produced? [38, 40.]

25. When 60 c.c. of a mixture of nitrous and nitric oxides were mixed with an equal volume of pure hydrogen, and exploded, 38 c.c. of pure nitrogen were left; find the quantity of each gas in the original mixture.

26. A mixture of marsh gas and hydrogen with excess of oxygen measures 67 c.c., but after explosion measures 23 c.c., of which 7 c.c. are not absorbed by potash; required the composition of the original mixture.

27. A mixture of 74 c.c. of oxides of carbon with 26 c.c. of oxygen left after explosion 88 c.c. of gas; required the quantities of each gas originally present.

28. 40 c.c. of a mixture of oxygen and nitrogen were mixed with 37 c.c. of hydrogen and exploded. The gas which remained measured 32 c.c.; what was the composition of the original mixture?

29. What volume of oxygen can be obtained from 73·08 grams of manganese dioxide—

(*a*) When heated alone,

(*b*) When heated with strong sulphuric acid? [3, 5.]

30. Find what mass of chlorine would be obtained by the action of excess of strong hydrochloric acid upon 290 grams of manganese dioxide, containing 70 per cent. of MnO_2. [44.]

31. A sample of manganese dioxide is found to contain 34 per cent. of oxygen. How much chlorine can be obtained by acting with strong hydrogen chloride on 20 grams of it? [44.]

32. What mass of potassium chlorate will yield sufficient oxygen to convert 10 grams of carbon into carbon dioxide, and what space will this amount of carbon dioxide occupy at 100° C. and 760 m.m.? [1, 29.]

33. A rectangular gasholder—

(1) 75 cm. long, 25 cm. broad, and 20 cm. deep, is to be

filled with hydrogen at 7° C. and standard pressure; how much zinc at least is necessary? [6.]

(2) 60 cm. by 30 cm. by 50 cm., is to be filled with oxygen at 21° C. and 760 m.m.; how much potassium chlorate is necessary? [1.]

(3) 24 cm. by 50 cm. by 30 cm., is to be filled with nitrogen monoxide at 15° C. and 760 m.m.; how much ammonium nitrate is theoretically sufficient? [14.]

34. A cylindrical gasholder ($\pi = \frac{22}{7}$)—

(1) 60 cm. long and 25 cm. in diameter, is to be filled with oxygen at 13° C. and 760 m.m.; how much potassium chlorate is required? [1.]

(2) 50 cm. long, and 17 cm. in radius, is to be filled with carbon dioxide at 16° C. and standard pressure; how much marble is necessary? [28.]

35. A closed room 8 metres long, 5 metres wide, and 4 metres high, is to be disinfected by burning in it carbon disulphide. What mass of liquid must be burnt that $\frac{1}{10}$ of the oxygen in the room, measured at 0° C., may be converted into sulphur dioxide. (Assume 21 per cent. oxygen by volume in air, and 11·2 litres = volume of 1 gram hydrogen.) [65.]

36. If air contain 23 per cent. of its mass of oxygen, how many lbs. of carbon must be burnt in order to remove all the oxygen from 500 lbs. of air? [29.]

37. 10 c.c. of air, measured at atmospheric pressure, are introduced into the vacuum of a barometer when the mercury stands at 76 cm. The air then occupies 15 c.c.; by how much has the mercury column been depressed?

38. A certain element forms with chlorine a volatile compound, containing 54·6 per cent. of chlorine; the vapour density of the compound is 130; the specific heat of the element 0·055. Deduce the atomic weight of the element.

LIST OF ELEMENTS.

Non-Metals.

Element.	Symbol.	Atomic Weight.	Element.	Symbol.	Atomic Weight.
Arsenic	As	75	Iodine	I	127
Boron	B	11	Nitrogen	N	14
Bromine	Br	80	Oxygen	O	16
Carbon	Ca	12	Phosphorus	P	31
Chlorine	Cl	35·5	Silicon	Si	28
Fluorine	F	19	Sulphur	S	32
Hydrogen	H	1			

Metals.

Element.	Symbol.	Atomic Weight.	Element.	Symbol.	Atomic Weight.
Aluminium	Al	27·3	Magnesium	Mg	24
Antimony (Stibium)	Sb	122	Manganese	Mn	55
Barium	Ba	137	Mercury (Hydrargyrus)	Hg	200
Bismuth	Bi	210	Nickel	Ni	59
Cadmium	Cd	112	Platinum	Pt	197
Calcium	Ca	40	Potassium (Kalium)	K	39
Chromium	Cr	52	Silver (Argentum)	Ag	108
Cobalt	Co	59	Sodium (Natrium)	Na	23
Copper (Cuprum)	Cu	63·5	Strontium	Sr	87·5
Gold (Aurum)	Au	197	Tin (Stannum)	Sn	118
Iron (Ferrum)	Fe	56	Zinc	Zn	65
Lead (Plumbum)	Pb	207			

The atomic weights here given are those used in the foregoing calculations. More accurate values will be found in complete lists of the elements.

EQUATIONS.

E give here the various equations representing the reactions icated in the preceding examples. The student is advised all cases to try to form the equation for himself before erring to it in the following list:—

(1) $2\,KClO_3 = 2\,KCl + 3\,O_2$.

(2) $2\,HgO = Hg_2 + O_2$.

(3) $3\,MnO_2 = Mn_3O_4 + O_2$.

(4) $Cu_2 + O_2 = 2\,CuO$.

(5) $2\,MnO_2 + 2\,H_2SO_4 = 2\,MnSO + 2\,H_2O + O_2$.

(6) $Zn + H_2SO_4 = ZnSO_4 + H_2$.

(7) $Fe + H_2SO_4 = FeSO_4 + H_2$.

(8) $Fe + 2\,HCl = FeCl_2 + H_2$.

(9) $Na_2 + 2\,H_2O = 2\,NaHO + H_2$.

(10) $K_2 + 2\,H_2O = 2\,KHO + H_2$.

(11) $3\,Fe + 4\,H_2O = Fe_3O_4 + 4\,H_2$.

(12) $KNO_3 + H_2SO_4 = HKSO_4 + HNO_3$.

(13) $HNO_3 + NH_3 = (H_4N)NO_3$.

(14) $(H_4N)NO_3 = 2\,H_2O + N_2O$.

(15) $5\,N_2O + P_2 = P_2O_5 + 5\,N_2$.

(16) $S_2 + 2\,O_2 = 2\,SO_2$.

(17) $2\,N_2O + S = 2\,N_2 + SO_2$.

(18) $N_2O + K_2 = K_2O + N_2$.

(19) $3\,Cu + 8\,HNO_3 = 3\,Cu(NO_3)_2 + 4\,H_2O + 2\,NO$.

(20) $3\,Zn + 8\,HNO_3 = 3\,Zn(NO_3)_2 + 4\,H_2O + 2\,NO$.

(21) $10\,NO + 2\,P_2 = 5\,N_2 + 2\,P_2O_5$.

(22) $O_2 + 4\,NO = 2\,N_2O_3$.

(23) $O_2 + 2\,NO = 2\,NO_2$.

(24) $2\,Pb(NO_3)_2 = 2\,PbO + O_2 + 4\,NO_2$.

(25) $2\,(H_4N)Cl + CaO = CaCl_2 + H_2O + 2\,NH_3$.

(26) $9\,HNO_3 + 4\,Zn = 4\,Zn(NO_3)_2 + 3\,H_2O + NH_3$.

(27) $2\,H_2 + O_2 = 2\,H_2O$.

(28) $CaCO_3 + 2\,HCl = CaCl_2 + H_2O + CO_2$.

(29) $C_2 + 2\,O_2 = 2\,CO_2$.

(30) $CO_2 + Ca(HO)_2 = CaCO_3 + H_2O$.

(31) $2\,NaHO + CO_2 = Na_2CO_3 + H_2O$.

(32) $2\,KHO + CO_2 = K_2CO_3 + H_2O$.

(33) $H_2C_2O_4 = H_2O + CO_2 + CO$.

(34) $2\,CO_2 + C_2 = 4\,CO$.
(35) $2\,CO + O_2 = 2\,CO_2$.
(36) $CH_2O_2 = H_2O + CO$.
(37) $Na\,C_2H_3O_2 + Na\,HO = Na_2CO_3 + CH_4$.
(38) $CH_4 + 2\,O_2 = 2\,H_2O + CO_2$.
(39) $C_2H_6O = C_2H_4 + H_2O$.
(40) $C_2H_4 + 3\,O_2 = 2\,CO_2 + 2\,H_2O$.
(41) $C_2H_4 + 2\,Cl_2 = C_2 + 4\,HCl$.
(42) $H_2 + C_2H_2 = C_2H_4$.
(43) $2\,NaCl + 2\,H_2SO_4 + MnO_2 = MnSO_4 + Na_2SO_4 + 2\,H_2O + Cl_2$.
(44) $MnO_2 + 4\,HCl = MnCl_2 + 2\,H_2O + Cl_2$.
(45) $2\,KI + Cl_2 = 2\,KCl + I_2$.
(46) $NaCl + H_2SO_4 = HNaSO_4 + HCl$.
(47) $KCl + H_2SO_4 = HKSO_4 + HCl$.
(48) $H_2 + Cl_2 = 2\,HCl$.
(49) $HCl + NH_3 = (H_4N)\,Cl$.
(50) $Cu + 2\,H_2SO_4 = CuSO_4 + 2\,H_2O + SO_2$.
(51) $Hg + 2\,H_2SO_4 = HgSO_4 + 2\,H_2O + SO_2$.
(52) $NO_2 + SO_2 + H_2O = H_2SO_4 + NO$.
(53) $FeS + H_2SO_4 = FeSO_4 + H_2S$.
(54) $Cu + S = CuS$.
(55) $2\,H_2S + 3\,O_2 = 2\,H_2O + 2\,SO_2 = 2\,H_2SO_3$.
(56) $P_4 + 3\,KHO + 3\,H_2O = 3\,KH_2PO_2 + PH_3$.
(57) $3\,Ca(HO)_2 + 6\,H_2O + 2\,P_4 = 3\,Ca(PH_2O_2)_2 + 2\,PH_3$.
(58) $Na(H_4N)HPO_4 = NaPO_3 + H_2O + NH_3$.
(59) $2\,Na_2HPO_4 = H_2O + Na_4P_2O_7$.
(60) $Fe + CuSO_4 = Cu + FeSO_4$.
(61) $Zn + Pb(C_2H_3O_2)_2 = Pb + Zn(C_2H_3O_2)_2$.
(62) $HgCl_2 + 2\,KI = HgI_2 + 2\,KCl$.
(63) $C + 2\,H_2SO_4 = CO_2 + 2\,SO_2 + 2\,H_2O$.
(64) $PH_3 + 2\,O_2 = H_3PO_4$.
(65) $CS_2 + 3\,O_2 = CO_2 + 2\,SO_2$.
(66) $HNO_3 + KHO = KNO_3 + H_2O$.
(67) $4\,NH_3 + 3\,O_2 = 2\,N_2 + 6\,H_2O$.
(68) $Mg_2 + CO_2 = C + 2\,MgO$.
(69) $2\,Cu(NO_3)_2 = 2\,CuO + O_2 + 4\,NO_2$.
(70) $4\,H_3PO_3 = 3\,H_3PO_4 + PH_3$.
(71) $PCl_5 + H_2O = 2\,HCl + POCl_3$.

ANSWERS

TO

NUMERICAL EXERCISES IN CHEMISTRY

WITH TABLES FURNISHING ADDITIONAL EXERCISES ON THE GASES

BY

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TABLES FURNISHING ADDITIONAL EXERCISES ON THE GASES.

The following tables show the masses of the various substances required to form the mass or volume of gas specified, and of the substances formed at the same time as the gas. They are intended to furnish materials for constructing additional examples on chemical equations, if such are required, by substitution of the numbers in them for those contained in the examples to Chapter IX.

OXYGEN.

Potassium Chlorate.	Potassium Chloride.	Oxygen (grams).	Oxygen (c.c.).
$9{\cdot}14\dot{6}$	$5{\cdot}562\dot{6}$	3·584	2500
$11{\cdot}4\dot{3}$	$6{\cdot}95\dot{3}$	4·48	3125
13·72	8·344	5·376	3750
16·464	10·0128	6·4512	4500
20·58	12·516	8·064	5625
34·3	20·86	13·44	9375

HYDROGEN.

Zinc.	Hydrogen Sulphate.	Zinc Sulphate.	Hydrogen (grams).	Hydrogen (c.c.).
7·28	10·976	18·032	·224	2500
9·1	13·72	22·54	·28	3125
10·92	16·464	27·048	·336	3750
13·104	19·7568	32·4576	·4032	4500
16·38	24·696	40·572	·504	5625
27·3	41·16	67·62	·84	9375

NITROGEN MONOXIDE.

Ammonium Nitrate.	Water.	Nitrogen Monoxide (grams).	Nitrogen Monoxide (c.c.).
8·96	4·032	4·928	2500
11·2	5·04	6·16	3125
13·44	6·048	7·392	3750
16·128	7·2576	8·8704	4500
20·16	9·072	11·088	5625
33·6	15·12	18·48	9375

NITROGEN DIOXIDE.

Copper.	Hydrogen Nitrate.	Copper Nitrate.	Water.	Nitrogen Dioxide (grams).	Nitrogen Dioxide (c.c.).
10·668	28·224	31·5	4·032	3·36	2500
13·335	35·28	39·375	5·04	4·2	3125
16·002	42·336	47·25	6·048	5·04	3750
19·2024	50·8032	56·7	7·2576	6·048	4500
24·003	63·504	70·875	9·072	7·56	5625
40·005	105·84	118·125	15·12	12·6	9375

AMMONIA.

Quicklime.	Ammonium Chloride.	Calcium Chloride.	Water.	Ammonia (grams).	Ammonia (c.c.).
3·136	5·992	6·216	1·008	1·904	2500
3·92	7·49	7·77	1·26	2·38	3125
4·704	8·988	9·324	1·512	2·856	3750
5·6448	10·7856	11·1888	1·8144	3·4272	4500
7·056	13·482	13·986	2·268	4·284	5625
11·76	22·47	23·31	3·78	7·14	9375

CARBON DIOXIDE.

Calcium Carbonate.	Hydrogen Chloride.	Water.	Calcium Chloride.	Carbon Dioxide (grams).	Carbon Dioxide (c.c.).
11·2	8·176	2·016	12·432	4·928	2500
14	10·22	2·52	15·54	6·16	3125
16·8	12·264	3·024	18·648	7·392	3750
20·16	14·7168	3·6288	22·3776	8·8704	4500
25·2	18·396	4·536	27·972	11·088	5625
42	30·66	7·56	46·62	18·48	9375

CARBON MONOXIDE.

Oxalic Acid.	Carbon Dioxide.	Water.	Carbon Monoxide (grams).	Carbon Monoxide (c.c.).
10·08	4·928	2·016	3·136	2500
12·6	6·16	2·52	3·92	3125
15·12	7·392	3·024	4·704	3750
18·144	8·8704	3·6288	5·6448	4500
22·68	11·088	4·536	7·056	5625
37·8	18·48	7·56	11·76	9375

MARSH GAS.

Sodium Acetate.	Caustic Soda.	Sodium Carbonate.	Marsh Gas (grams).	Marsh Gas. (c.c.).
9·184	4·48	11·872	1·792	2500
11·48	5·6	14·84	2·24	3125
13·776	6·72	17·808	2·688	3750
16·5312	8·064	21·3696	3·2256	4500
20·664	10·08	26·712	4·032	5625
34·44	16·8	44·52	6·72	9375

OLEFIANT GAS.

Alcohol.	Water.	Olefiant Gas (grams).	Olefiant Gas (c.c.).
5·152	2·016	3·136	2500
6·44	2·52	3·92	3125
7·728	3·024	4·704	3750
9·2736	3·6288	5·6448	4500
11·592	4·536	7·056	5625
19·32	7·56	11·76	9375

CHLORINE.

Sodium Chloride.	Hydrogen Sulphate.	Manganese Dioxide.	Sodium Sulphate.	Manganese Sulphate.	Water.	Chlorine (grams).	Chlorine (c.c.).
13·104	21·952	9·744	15·904	16·912	4·032	7·952	2500
16 38	27·44	12·18	19·88	21·14	5·04	9·94	3125
19·656	32·928	14·616	23·856	25·368	6·048	11·928	3750
23 5872	39·5136	17·5392	28·6272	30·4416	7·2576	14·3136	4500
29·484	49·392	21 924	35·784	38·052	9·072	17·892	5625
49 14	82·32	36·54	59·64	63·42	15·12	29·82	9375

HYDROGEN CHLORIDE.

Sodium Chloride.	Hydrogen Sulphate.	Hydrogen Sodium Sulphate.	Hydrogen Chloride (grams).	Hydrogen Chloride (c.c.)
6·552	10·976	13·44	4·088	2500
8·19	13·72	16·8	5·11	3125
9·828	16·464	20·16	6·132	3750
11·7936	19·7568	24·192	7·3584	4500
14·742	24·696	30·24	9·198	5625
24·57	41·16	50·4	15·33	9375

SULPHUR DIOXIDE.

Copper.	Hydrogen Sulphate.	Copper Sulphate.	Water.	Sulphur Dioxide (grams).	Sulphur Dioxide (c.c.).
7·112	21·952	17·864	4·032	7·168	2500
8·89	27·44	22·33	5·04	8·96	3125
10·668	32·928	26·796	6·048	10·752	3750
12·8016	39·5136	32·1552	7·2576	12·9024	4500
16·002	49·392	40·194	9·072	16·128	5625
26·67	82·32	66·99	15·12	26·88	9375

HYDROGEN SULPHIDE.

Iron Sulphide.	Hydrogen Sulphate.	Iron Sulphate.	Hydrogen Sulphide (grams).	Hydrogen Sulphide (c.c.).
9·856	10·976	17·024	3·808	2500
12·32	13·72	21·28	4·76	3125
14·784	16·464	25·536	5·712	3750
17·7408	19·7568	30·6432	6·8544	4500
22·176	24·696	38·304	8 568	·5625
36·96	41·16	63·84	14·28	9375

ANSWERS.

I.

1. (1) 7000. (2) 0·5. (3) 300,000. (4) 330. (5) 20,500. (6) 110·3. (7) 3740. (8) 4200. (9) 353·9.
2. (1) 0·34. (2) 0·106. (3) 8000·9. (4) 420·04. (5) 27. (6) 0·005361.
3. (1) 7. (2) 0·064. (3) 8000. (4) 0·0095. (5) 5000·005. (6) 3000. (7) 5·67. (8) 2002. (9) 0·0395.
4. (1) 870·4. (2) 42040. (3) 600606. (4) 1211·1. (5) 18·3.
5. (1) 20·391. (2) 17·632. (3) 7·654.
6. 12·4 cm. 7. 20. 8. 150. 9. 0·08 litre. 10. 2 cm.
11. 90. 12. 60 metres.
13. 5, 10, 15, . . . inches with 127, 254, . . . m.m. 14. 16·09.
15. 40·08. 16. 32·2.

II.

1. (1) 59° F., 12° R. (2) 95°, 28°. (3) 122°, 40°. (4) —112°, —64°. (5) —49°, —36°. (6) 70°·7, 17°·2. (7) 92°·3, 26°·8. (8) 20°·3, —5°·2. (9) —22°·9, —24°·4. (10) —459°·22, —218°·32.
2. (1) 10° C., 8° R. (2) 20°, 16°. (3) 25°, 20°. (4) —20°, —16°. (5) —30°, —24°. (6) 18°, 14°·4. (7) 4°, 3°·2. (8) —21°, —16°·8. (9) —28°, —22°·4. (10) —34°, —27°·2.
3. (1) 86° F., 30° R. (2) 158°, 70°. (3) 194°, 90°. (4) 14°, —10°. (5) —67°, —55°. (6) 73°·4, 23°. (7) 142°·7, 61°·5, (8) 25°·7, —3°·5. (9) —156°·1, —104°·5. (10) —227°·2, —144°.
4. (1) 20° C., 16° R. (2) 45°, 36°. (3) 25°, 20°. (4) $18\frac{1}{3}$°, $14\frac{2}{3}$°. (5) 16°, 12°·8.
5. (1) 45° F., 20° R. (2) 72°, 32°. (3) 18°, 8°. (4) 99°, 44°. (5) 110°·7, 49°·2.
6. 27°. 7. (*a*) —40°. (*b*) —25°·6. (*c*) 0°.
8. All those of the form $(9m + 32)$° F, $5m$° C, $4m$° R., where m is an integer; 19.
9. 1 : 1·8 : 2·25; 0·16, 0·288, 0·36 cm. 13. 68° F., 20° C.
14. 14° F., —10° C. 15. 41° F., 5° C. 16. —4° F., —20° C.
17. 39°·2 F., 4° C. 20. 27·432. 21. 964·8.

III.

1. (1) 5. (2) 1·1. (3) 2·4. 2. (1) 13 c.c. (2) 7 c.c. (3) 11·3 c.c.
3. (1) 77 grams. (2) 8·82 grams. (3) 5·64 grams.
4. (1) 3924 dynes. (2) 5886 dynes. (3) 2648·7 dynes.
5. 1·2 grams. 6. 67·98 grams. 7. 15 c.c. 8. 10 c.c. 10. 3·4.
11. 33·99. 12. 999999·84. 13. 49 : 26. 14. 0·875.
15. 6·71. 16. 453·2 m.m. 17. 1·25. 18. 5. 19. 9·0594 . . .
20. 0·8.

IV.

1. 153 c.c. 2. 708 m.m. 3. 5 c.c. 4. 185 m.m.
5. 65 m.m. 6. 18 c.c. 7. 756 m.m. 8. 772 m.m.
9. 76·5 c.c. 10. Reduced two-thirds. 11. 64·6 m.m.
12. 14·7. 13. 665 m.m. 14. Pressure doubled, volume halved.
15. 152 m.m. increase. 17. 5 cm. 18. 10 m.m. 19. 570 m.m.
20. 168 m.m., 588 m.m.

V.

1. 77 c.c. 2. 432 m.m. 3. 148 c.c. 4. 6°.
5. 251 m.m. 6. 30 c.c. 7. 48^{c}·6. 8. 517 m.m.
10. 9 m.m., 2 c.c. increase. 12. 27·5625 litres. 13. 533 grains.
14. 84° C. 15. 15·1$\dot{6}$, 30·$\dot{3}$, 7·58$\dot{3}$, 14 21875.
16. 1·4336, 1·2544, 1·9712, 1·344, 0·7616, 1 2544, 1·9712, 0·7168, 3·1808,
2·8672; 1·3589$\dot{3}$, 1·1890$\dot{6}$, 1·8685$\dot{3}$, 1·274, 0·7219$\dot{3}$, 1·1890$\dot{6}$, 1·8685$\dot{3}$,
0·6794$\dot{6}$, 3·0151$\dot{3}$, 2·7178$\dot{6}$ grams.

VI.

1. 1 c.c. 2. 1 : 16. 3. 5 grams. 4. 2 : 1. 5. Equal.
6. 1·25 litre.

VII.

1. 20°. 2. 21°. 3. 6 grams. 4. 7°. 5. 24°. 6. 0·11.
7. 30°. 8. 56°. 9. 40°. 10. 0·095. 11. 0·05.
12. 0·0308. 13. 0·05.

VIII.

1. 10°. 2. 5 grams. 3. 3 grams. 4. 13°. 5. 36°.
6. 89·5 grams of water at 0° C. 7. 80. 8. 0·1808 lb. 9. 535.
10. 1565 grams. 11. 124°. 12. 5·37. 13. 144.

IX.

Oxygen.

1. (1) 490, 298 grms. (2) 44·1, 26·82 oz. (3) 53·9, 32·78 grms. (4) 73·5, 44·7 grms. (5) 188·65, 114·73 grms.
2. (1) 54·88. (2) 6·86. (3) 8·232. (4) 201·226̇ grms.
3. (1) 27 grms. (2) 94·5 grms. (3) 96·768 grms. (4) 35·1 oz. (5) 120·96 grms.
4. (1) 8 lbs. (2) 24 grms. (3) 22·4 grms. (4) 35·2 oz. (5) 67·2 grms.
5. (1) 600 grms. (2) 74·5. (3) 343·5 grms.
6. 781·25. 7. 20·2176 grms. 8. 32 grms. 9. 31·8 grms.
10. 1375 c.c., 0·7392 grms.

Hydrogen.

1. (1) 195, 294, 483 grms. (2) 162·5, 245, 402·5 grms. (3) 52, 78·4, 128·8 oz. (4) 117, 176·4, 289·8 grms. (5) 136·5, 205·8, 338·1 oz.
2. (1) 29·12 grms. (2) 43·68. (3) 1·82. (4) 61·152. (5) 9·1728.
3. (1) 1·4, 68·6 grms. (2) 2·2, 107·8. (3) 11·2, 548·8. (4) 7, 343. (5) 9·8, 480·2.
4. (1) 224, 292 grms. (2) 50·4, 65·7 oz. (3) 62·72, 81·76 grms. (4) 156·8, 204·4 grms. (5) 7·84, 10·22 grms.
5. 1406·25 c.c. 6. 160, 48 grms. 7. 14·43 q.p. 8. 165·62 grms. 9. 745·29 grms. 10. 98·784 grms.

Nitric Acid.

1. (1) $14\frac{3}{7}$, 14 grms. (2) 89·7̇, 87·1̇ grms. (3) 33·6̇, 32·6̇ oz. (4) 25·25, 24·5 grms. (5) 20·2, 19·6 grms.
2. (1) 9·45, 20·4 grms. (2) 69·3, 149·6 grms. (3) 44·1, 95·2 oz. (4) 132·3, 285·6 grms. (5) 176·4, 380·8 grms.
3. 42, 7 grms. 4. 92·983̇%. 5. 27 grms.
6. (1) 37·8. (2) 85. (3) 331. (4) 66 grms.
7. 85 grms. 8. 108.

Nitrogen Monoxide.

1. (1) 33 grms. (2) 55 grms. (3) 14·3 oz. (4) 7·7 grms. (5) 12·1 grms.
2. (1) 35·84 grms. (2) 2·24. (3) 53·76. (4) 12·096. (5) 24·64.
3. 795·2 grms, 625 litres. 4. 6 litres. 5. 13·16 grms.
6. 76·44 grms. 7. 150 litres. 8. 218·4 c.c.
9. 6·97, 6·41 litres q.p. 10. 8·0388 grms.

Nitrogen Dioxide.

1. (1) 38·1, 100·8, 112·5 grms. (2) 57·15, 151·2, 168·75 oz. (3) 22·86, 60·48, 67·5 grms. (4) 41·91, 110·88, 123·75 grms. (5) 34·29, 90·72, 101·25 grms.
2. (1) 42·672 grms. (2) 26·67. (3) 149·352. (4) 18·669. (5) 58·674.
3. 225 grms, 83·705 litres q.p. 4. 19 litres. 5. 68·796 grms.
6. 56·007. 7. 1 : 1·008. 8. 11 : 15. 9. 42·56 grms.
10. 6250 c.c.

Ammonia.

1. (1) 28, 53·5, 55·5 grms. (2) 8·4, 16·05, 16·65 oz. (3) 70, 133·75,

138·75 grms. (4) 30·8, 58·85, 61·05 oz. (5) 58·8, 112·35, 116·55 grms.
2. (1) 1250 c.c. (2) 10 litres. (3) 6250 c.c. (4) 8750 c.c. (5) 55 litres.
3. 6·048 grms. 4. 9375. 5. 1520 m.m. 6. 109·2 grms.
7. 250 litres. 8. 130 grms. 9. 683·655 grms. 10. 25%.

Carbon Dioxide.

1. (1) 75, 54·75, 83·25 grms. (2) 35, 25·55, 38·85 grms. (3) 60, 43·8, 66·6 grms. (4) 140, 102·2, 155·4 grms. (5) 220, 160·6, 244·2 grms.
2. (1) 4 litres. (2) 875 c.c. (3) 6 litres. (4) 2800 c.c. (5) 15 litres.
3. 13·44, 49·92 grms. 4. 1·5 grm. 5. 33 grms. 6. 517 cwt.
7. 176·4 grms. 8. 19 litres. 9. 140·14 grms. 10. No; 0·57 . . . %.

Carbon Monoxide.

1. (1) 3·08 grms. (2) 7 oz. (3) 4·2 grms. (4) 9·8 grms. (5) 7·7 grms.
2. (1) 20 litres. (2) 35 litres. (3) 1375 c.c. (4) 6875 c.c. (5) 875 c.c.
3. 56 grms. 4. 15 grms. 5. 33 grms. 6. 69 grms.
7. 20 litres. 8. 687·96 grms. 9. 10 c.c. CO, 5 c.c. CO_2.
10. 50 litres.

Marsh Gas.

1. (1) 41, 20, 53 grms. (2) 14·35, 7, 18·55 grms. (3) 22·55, 11, 29·15 grms. (4) 26·65, 13, 34·45 grms. (5) 73·8, 36, 95·4 grms.
2. (1) 20 litres. (2) 625 c.c. (3) 35 litres. (4) 15 litres.
3. 655·2 grms. 4. 250 litres. 5. 25%. 6. 436·8 grms.
7. 45 grms H_2O, 55 grms. CO_2. 8. 116·2 cub. ft. 9. 7 c.c.
10. 904,000 cub. ft.

Olefiant Gas.

1. (1) 69 grms. (2) 80·5 grms. (3) 27·6 oz. (4) 25·3 grms. (5) 149·5 grms.
2. (1) 8·064 grms. (2) 2·2176. (3) 44·352. (4) 52·416 grms.
3. (*a*) 66 grms CO_2, 27 grms. H_2O. (*b*) 18 grms. C, 109·5 grms. HCl
4. 5 c.c. O, 10 c.c. CO_2. 5. (*a*) Equal. (*b*) 1·75 : 1. 6. 21·504 grms.
7. 41·16 grms. 8. 37½ litres. 9. 392 grms.
10. 2 c.c. CH_4, 1 c.c. C_2H_4.

Chlorine.

1. (1) 46·8, 78·4, 34·8 grms. (2) 70·2, 117·6, 52·2 grms. (3) 105·3, 176·4, 78·3 oz. (4) 128·7, 215·6, 95·7 grms. (5) 163·8, 274·4, 121·8 grms. (6) 655·2, 1097·6, 487·2 grms. (7) 32·76, 54·88, 24·36 grms. (8) 78·624, 131·712, 58·464 grms.
2. (1) 113·6, 120·8 grms. (2) 27·832, 29·596. (3) 184·6, 196·3.
3. 3572·933 grms.
4. (1) 135·5 grms. (2) 269. (3) 149·75.

2·47. 6. 14·2 grms. 7. 137·592 grms. 8. 19 litres.
1·524 grm. 10. 14·2 grms.

Hydrogen Chloride.

1. (1) 117, 196, 240 grms. (2) 10·53, 17·64, 21·6 oz. (3) 24·57, 41·16, 50·4 grms. (4) 128·7, 215·6, 264 grms. (5) 175·5, 294, 360 grms.
2. (1) 26·208, 43·904 grms. (2) 1·638, 2·744. (3) 327·6, 548·8. (4) 40·95, 68·6.
3. 204·4 grms.
4. (1) 94·5 grms. (2) 149. (3) 269. (4) 107.
5. 6·78 grms. 6. 3 : 3·4. 7. 75 litres. 8. 114·66 grms.
9. 30·3 per cent. q.p. 10. 255·5 grms.

Sulphur.

1. (1) 15·875, 49, 39·875 grms. (2) 17·78, 54·88, 44·66. (3) 25·4, 78·4, 63·8. (4) 69·85, 215·6, 175·45. (5) 76·2, 235·2, 191·4.
2. (1) 5 litres. (2) 6250 c.c. (3) 8750 c.c. (4) 35 litres. (5) 60 litres.
3. 96 grms. 4. 147 grms. 5. 4 grms. 6. 91, 89·18 grms.

Hydrogen Sulphide.

1. (1) 26·4, 29·4, 45·6 oz. (2) 29·04, 32·34, 50·16 grms. (3) 96·8, 107·8, 167·2 grms. (4) 184·8, 205·8, 319·2 grms. (5) 343·2, 382·2, 592·8 grms.
2. (1) 19·712, 21·952 grms. (2) 7·8848, 8·7808. (3) 24·64, 27·44.
3. (1) 47·75. (2) 116. (3) 86 grms. 4. 68 grms.
5. 27 grms. H_2O, 96 grms. SO_3. 6. 7·168 grms. 7. 22·0416 grms.
8. 98·784 grms. 9. 57 litres. 10. 4484·48 grms.

Phosphorus.

1. 17 lbs. 2. 17 grms. 3. 12·75 grms.
4. (1) 6·2 grms. (2) 15·5 grms. (3) 7·75 grms.
5. 30·6 grms. 6. 99·4 grms. 7. 49 : 41. 8. 31¼ litres.
9. 1562½ c.c. 10. 61·4 grms.

X.

1. 5, 40. 2. 20 grms. 3. 24 grms. 4. 20 c.c. 5. 91 c.c.
6. 46·8, 41·6 grms. 7. 63, 7 grms. 8. 31·25.
9. 115, 155 grms. 10. 1·4 : 1. 11. 3 grms., 27 grms.
12. (1) 2, 58 oz. (2) 2·6, 75·4 grms. (3) 1·4, 40·6 lbs. (4) 8·4, 243·6 grms.
13. 20, 3·6 grms. 14. 98° C. 15. 43·227 m.m. 16. 99° C.
17. 1876·4 metres. 18. 98°·1 C.

XI.

1. 3 litres. 2. (*a*) 250·88, 71·68 grms. (*b*) 244·608, 69·888 grms.
3. 29·792, 8·512 grms. 4. 20 litres. 5. 40 litres. 6. 8·904 grms.
7. 250 litres. 8. 63·616 grms. 9. 4 grms. 10. 85·5 grms.
11. (1) 4·89 grms. (2) 10·009. (3) 50·922. (4) 3·1268 grms. q.p.
12. 362·95 grms. q.p. 13. 0·3086 q.p. 14. 374·1 c.c. q.p.
15. 8·183 grains q.p.

XII.

1. (1) 17. (2) 11. (3) 142. (4) 12.
2. (1) 273. (2) 20. (3) 21.
3. 9. 4. 21. 5. 6 litres. 6. 0·6875 litre. 7. 15·68
8. 500.

XIII.

1. 308·8 ft. 2. 0·25° C. 3. 366·7 ft. 4. 50° F.
5. 8118 calories. 6. (i) 103386. (ii) 5650. (iii) 3692 cal.
7. 5 grms. 8. 3829⅓ grms. 9. 4·265 ; 1 q.p.
10. 414 calories q.p.
11. (*a*) 300,000 *g* ergs. (*b*) 112 *g* ergs. (*c*) 134,400 foot-pounds.
12. 3,125,000,000 ergs.

XIV.

1. (1) 40 Ca, 12 C, 48 O. (2) 50 S, 50 O. (3) 60 Mg, 40 O. (4) 20 Mg,
26·$\dot{6}$ S, 53·$\dot{3}$ O. (5) 57·5 Na, 2·5 H, 40 O. (6) 5 H, 35 N, 60 O.
(7) 8·$\dot{3}$ H, 29·1$\dot{6}$ N, 12·5 C, 50 O. (8) 52 Zn, 9·6 C, 38·4 O.
(9) 2·$\dot{2}$ H, 26·$\dot{6}$ C, 71·$\dot{1}$ O. (10) 0·8$\dot{3}$ H, 19·1$\dot{6}$ Na, 26·$\dot{6}$ S, 53·$\dot{3}$ O.
2. (1) $Na_3Al\ F_6$. (2) Fe As S. (3) $Na_4P_2O_7$. (4) $Na\ P\ O_3$.
(5) $Cu_2S + Fe_2S_3$. (6) $Fe\ Cr_2O_4$. (7) $Na_2B_4O_7$. (8) $C\ H\ Cl_3$.
(9) $C\ O\ (N\ H_2)_2$. (10) $K_3Fe\ (C\ N)_6$.
3. $Ca\,C\,O_3$. 4. C_2H_6O. 5. $H_8C_8O_2$. 6. 9·12 grms.
7. 0·033 q.p. 8. 19·2 grms. of iron q.p. 9. 10°·5 C. 10. 105°.
11. 56·239 q.p. 12. 56. 13. 290 : 571.
14. (1) 50·685, 49·315. (2) 48·78, 51·22. (3) 56·1, 43·9. (4) 63·$\dot{6}\dot{3}$,
36·$\dot{3}\dot{6}$. (5) 63·82, 36·18 grms.
15. 62 lbs. 16. 202·5 grms. 17. 2 grms. 18. 21·2 grms
19. 67·75 grms. 20. 1·095 grm.
21. 250 litres $S\ O_2$, 125 litres $C\ O_2$.
22. (1) 12 c.c. N, 36 c.c. H. (2) 25 c.c. N, 25 c.c. O. (3) 6 litres N,
15 O. (4) 2500 c.c. H, 2500 Cl. (5) 100 c.c. N, 50 O.
23. (1) ½ litre O, 1 $C\ O_2$. (2) 3 O, 2 $C\ O_2$. (3) 1½ O, 1 $S\ O_2$. (4) 2
litres O.
24. 80 c.c. O, 50 $C\ O_2$. 25. 16 c.c. N_2O, 44 N O.
26. 16 c.c. $C\,H_4$, 8 H, 43 O. 27. 24 c.c. C O, 50 $C\ O_2$.
28. 15 c.c. O, 25 N. 29. 6250 c.c., 9375 c.c. 30. 165·6 grms.
31. 21·25 grms. 32. 68·0$\dot{5}$ grms., 25·415 litres q.p.
33. (1) 106·47 grms. (2) 305·76 grms. (3) 611·52 grms.
34. (1) 102·9 grms. (2) 192·192 grms.
35. 5700 grms. 36. 43·125 lbs. 37. 50⅔ cm. 38. 118.